EXPERIMENTAL EDITION

SCHOOL CHEMISTRY

CLASS 7

TEACHER'S GUIDE

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RESEARCH AND TRAINING

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SCHOOL CHEMISTRY

CLASS 7

TEACHER'S GUIDE

Prepared by

NCERT Chemistry Study Groups

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NEW DELHI

PREFACE

From time to time new textbooks have been written in chemistry, mainly to present additional facts brought to light by researches. In almost all these books the facts are given as a series of statements It is usual to write, for instance, that "acids are substances having a sour taste, are corrosive to the skin, turn blue litmus red, contain hydrogen which can be liberated by metals like zinc iron, etc".

The pupils read and memorise these 'important' facts and practically repeat the statements in an examination It is only at a much later stage, if at all, that they may be called to verify some of the statements in the laboratory. The thrill and joy of discovery is lacking, as they have already been told what to expect. Such a study becomes dull and uninteresting and is hardly different from a study of remote events in history. This is how most of us may have learnt science.

Contrast this approach with the one where the pupils first do things by themselves in the laboratory, record their observations, draw conclusions and then try to find some reasonable explanation for the observed facts. For instance, the pupils may be given a few unknown solutions Some of them are acids, some bases and the rest neither The pupils are advised to try the action of these solutions on litmus, iron filings and sodium carbonate, etc. They record their observations and find for themselves, when told what the solutions are, the characteristic properties of acids and bases and how these differ from one another and from others. Learnt this way, knowledge is more easily retained, without the strain of memorization. There is also the thrill of finding something new by themselves.

This is the new approach which lays emphasis on pupils' experimentation and learning through their own observations. The topics covered in the text follow, and not precede, the experiments in the laboratory manual. The teacher's role is mainly to guide in laying down the general procedure and help in organising the laboratory work. No statements or facts are expected to be made which the pupils are supposed to accept and remember. Doubtless, certain topics cannot be experimentally verified at such an early stage by the pupils. Such topics may be explained with the aid of charts, diagrams and models and, wherever available, by means of filmstrips. However, topics of a theoritical nature are discussed only to the extent necessary for an intelligent understanding and correlation of the observed facts.

Teaching chemistry according to this new approach requires considerable effort on the part of the teacher. The members of the Study Groups hope that the teacher's guide will prove to be of some help to the teacher in achieving this objective. Though the teacher is expected to keep to the scheduling of the units, he has considerable freedom in organising the detailed work. Although certain experiments have been suggested, he may design new experiments to fulfil the same objectives. Without experimental basis the course will not be meaningful.

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TO THE TEACHER

In view of the great strides that chemistry has made in recent years, it has become necessary to expose the pupils to the modern ideas and trends in chemistry (even at the school level) so that they keep pace with modern developments. To make the study of the subject interesting and intelligible, a new approach is attempted in which the pupils are made to learn the basic principles of chemistry through experimentation instead of the traditional way of teaching which leads to memorisation of facts. This approach is believed to promote a sense of experimentation and inquiry in the young minds. Besides, it enables them to develop sufficient skill in the laboratory techniques and appreciate the need for careful observation.

May be it is the first time a project of this type is placed before you for teaching. The success of this depends on your effort. The facts learnt in the laboratory and through demonstration experiments should form the subject matter of the text Before each laboratory session, careful planning and preparation is necessary for successful conduct of the practical class. Gather all the materials required (apparatus, chemicals and solutions) well before the class commences. The materials required for each experiment are mentioned in the laboratory manual. Supply the pupils only the required amounts of the materials to avoid wastage. A list of apparatus and the amounts of chemicals required for a class of 40 pupils for all the experiments of the class is given in this guide so that the laboratory is equipped with all the required chemicals and apparatus.

It is desirable that, as far as possible, the experiments are done by the pupils individually. Try to use, wherever possible, improvised apparatus, made of materials locally available. The pupils should be taught the correct way of handling the materials. It is important that they are not told as to what will happen before hand. This would spoil their keen observation, curiosity, and spirit of inquiry. Go round the class to see whether the pupils are doing the experiments properly and check if they missed any observation (s). See that the pupils maintain a separate note book for recording observations. Check the answers of the pupils soon after the completion of each experiment and make them repeat the experiment if the observations are incorrect. (Answers to questions under each experiment are given in this guide for your guidance).

To supplement the experimental work and to cover as many concepts of the syllabus as possible, some demonstrations are described in this guide. The equipment and chemicals needed for them are also listed separately They should be procured and the demonstrations performed well before they are shown in the class. Under each demonstration, a few questions are suggested. Many such, questions may be asked during the conduct of the demonstration to direct the pupils attention to important points to be taken note of.

Follow as far as possible the schedule given in each unit. Needless to say it is very useful for planning the lesson. One lesson consists of two consecutive periods of forty minutes duration each. At the end of each unit a quiz is given. It is only suggestive. You may prepare many such quizes for examining the pupils.

It is hoped that the instruction material of this guide will be helpful for successful handling of the class. The background information is meant for extending the horizons of your knowledge and should not be taught to the pupils The teacher's guide is meant for your use and should not be placed in the hands of the pupils Any helpful suggestions and comments are welcome.

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5

Chemical Reactions

INSTRUCTION MATERIAL

Introduction

Important facts about the nature and behaviour of substances are revealed during a study of chemical reactions. The general characteristics of chemical reactions, conditions necessary for reactions to occur, the factors affecting the rate of a reaction and the main types of reactions are discussed in this unit.

Outline

1. What is a chemical reaction? (5.1).
2. Characteristics of a chemical reaction (5.2).
3. Conditions that help bring about a chemical reaction (5.3).
4. Factors which influence the rate of a reaction (5.4).
5. Types of chemical reactions (5.5).

Concepts

1. A chemical reaction is said to have taken place when a given substance(s) changes to a substance(s) having different properties.
2. Chemical reactions are accompanied by energy changes — usually in the form of heat and occasionally in the form of light and electrical energy. In some reactions, however, energy changes are too small to be measured.
3. For a chemical reaction to occur, it is necessary that the reactants be brought into intimate contact and usually some form of energy like heat, light or electricity be supplied.

4. The rate of a chemical reaction is influenced by the particle size of the reactants, their concentrations, temperature and the presence of a catalyst.

5. Reactions may be classified into a number of types such as,

   a) Combination b) Decomposition
   c) Displacement d) Double decomposition

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	17	Discussion	—	5.1
2	18	—	I and II	5.2
3	—	—	III, IV, V and VI	5.3
4	19	—	VII	5.4
5	—	—	VIII	5.4 and 5 5
6	—	—	—	Review and quiz

Development

5.1 What is a chemical reaction ?

Do not tell the pupils what constitutes a chemical reaction. Ask them to carry out carefully the reactions in experiments 17 and 18 and record their observations in the tabular form given in the laboratory manual.

Experiment 17 Study of chemical reactions-I

Time required . 40 minutes

Reaction I Burning of magnesium

Tell the pupils that commercially available magnesium ribbon (or wire) is coated with an oxide film. The film is removed by rubbing with sand paper or more conveniently by dipping the magnesium ribbon in dilute hydrochloric acid (0.5 M) for a few seconds, washing it with water and wiping dry.

The light emitted by burning magnesium is rich in ultraviolet rays. These rays are harmful to the eyes and so, prolonged watching should be avoided Use of safety glasses is advised.

Reaction 2 Heating copper carbonate

The green copper carbonate decomposes to give black cupric oxide and carbon dioxide gas which turns freshly prepared lime water milky. The milky appearance is due to the formation of a precipitate of calcium carbonate.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3\downarrow + H_2O$$

Preparation of lime water

Obtain a good quality quick lime (about 50 g) and add it to about 500 ml of water. Filter the supernatant liquid through a funnel using cotton, just before the experiment.

Questions and answers

1. Why is it necessary to remove the delivery tube from the lime water before heating is stopped?
   To prevent back suction.
2. What gas is evolved in the reaction?
   Carbon dioxide.

Reaction 3 Heating mercuric oxide

Mercury and its compounds are poisonous Ask the pupils to wash their hands after the experiment.

Reaction 4 Reaction of iron with copper sulphate solution

i) Prepare a saturated solution of copper sulphate in water.
ii) See that the nails supplied to the pupils are clean.

Postlab discussion

Check the observations of the pupils, discuss each reaction and point out that in each case a new substance is formed. Avoid classification of the reactions into various types at this stage. Let the pupils make the classification at a later stage. Do not discuss as to why a chemical reaction takes place.

Explain that the magnesium ribbon starts burning only after it is heated in the flame to a certain high temperature. Once it starts burning the reaction (combination of magnesium with oxygen) liberates a large amount of heat, which is sufficient to maintain the rest of the ribbon at the necessary temperature even after the ribbon is removed from the flame.

Experiment 18 Study of chemical reactions-II

Time required : 40 minutes

Reaction 5 Addition of water to quicklime (calcium oxide)

Reaction 6 Mixing solutions of sodium sulphate and barium chloride

5% solutions of sodium sulphate (Na_2SO_4, $10H_2O$) and barium chloride ($BaCl_2$, $2\ H_2O$) should be supplied.

Reaction 7 Mixing solutions of copper sulphate and sodium carbonate

A 5% solution of copper sulphate ($CuSO_4$, $5\ H_2O$) and 2% solution of sodium carbonate (Na_2CO_3, $10\ H_2O$) should be supplied.

Reaction 8 Mixing ammonium chloride and sodium carbonate

Ensure that the pupils use dry samples of these substances.

No further postlab discussion is required. Merely mention that as in experiment 17, new substances are formed as a result of chemical reactions.

Correct observations for each reaction are given in the following table.

Reaction	Change in appearance of reactants	Evolution of heat	Production of light	Evolution of gas	Formation of precipitate	Any other observation
1 Burning of magnesium	Grey to white ash	✓	✓	x	x	x
2 $CuCO_3$ Heat	Green to black	x	x	✓	x	Gas turns lime water milky
3 HgO Heat	Turns black, becomes red on cooling	x	x	✓	x	Glowing splinter rekindled Silvery white globules deposit on the walls of the test tube
4 Fe+$CuSO_4$(aq)	Reddish-brown coating on iron nail	x	x	x	x	Blue colour of the solution fades gradually
5. CaO+H_2O	Lump turns to paste	✓	x	x	x	—
6 Na_2SO_4(aq)+ $BaCl_2$(aq)	White solid	x	x	x	✓	—
7. $CuSO_4$(aq)+ Na_2CO_3(aq)	Bluish green mass	x	x	x	✓	—
8 NH_4Cl + Na_2CO_3+ a drop or two of water	x	x	z	x	x	Smell of ammonia preceived

5.2 Characteristics of a chemical reaction

Discuss the various reactions in experiments 17 and 18 and elicit from the pupils that besides the formation of substances with different properties (from those of reactants) chemical reactions are characterised by energy changes. Show demonstrations I and II at this stage.

Demonstration I Chemical reaction is accompanied by the evolution of energy as heat and light

Time required : 10 minutes

Materials required

Asbestos square		Potassium chlorate	1 g
Mortar and pestle		Sugar	2 g
Dropper		Conc. sulphuric acid	0.5 ml
Crucible			

Procedure

Take about 1 g of powdered potassium chlorate and mix it with 2 g of powdered sugar (the two substances should be powdered separately). Place the mixture in a crucible kept on an asbestos sheet. Add three to four drops of concentrated sulphuric acid and stand aside. You will find that the mixture burns producing light and liberating much heat. White fumes are given off and a black mass is left behind.

Note : Avoid using materials in excess of the quantities suggested, as there may be risk of explosion.

Demonstration II Chemical reaction may also produce electrical energy

Time required : 15 minutes

Materials required

Beaker (250 ml)		Dilute sulphuric acid (3 M)	200 ml
Connecting wire		Copper strip (12 cm x 2 cm x 0.2 cm)	
Bulb 2.5V		Magnesium ribbon	15 cm
Bulb holder			
Crocodile clips	2		

Procedure

Take about 200 ml of dilute sulphuric acid in a beaker. Place a copper strip in it. Connect the strip by means of wire to a 2.5 volt bulb fixed in a holder. Connect the magnesium ribbon also to the bulb (Figure 5.1).

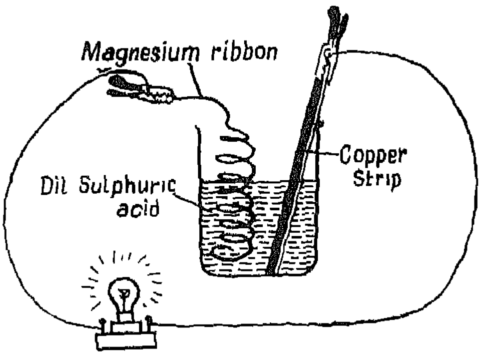


Figure 5 1 *Chemical reaction produces electrical energy*

Dip the magnesium ribbon in the acid. Immediate reaction sets in and the bulb glows. Gas is evolved at the magnesium (Do not keep the ribbon in the acid for a long time. Remove it after making sure that all the pupils have seen the bulb glow).

Discussion

When the circuit is completed, magnesium atoms from the ribbon enter the solution as magnesium ions, Mg^{++}. The electrons released move across the external circuit from the magnesium ribbon, through the filament of the bulb, to the copper strip. The passage of electrons through the filament (which has a high resistance) generates a large quantity of heat. Consequently the filament becomes red hot and the bulb glows.

5.3 Conditions that help bring about a chemical reaction

Demonstrations III to VI help the pupils to know some of the conditions that help bring about a chemical reaction.

Demonstration III Intimate contact between reactants is necessary for a reaction to take place

Time required . 10 minutes

Materials required

Mortar and pestle		Mercury	6 g (about 0.5 ml)
Test tubes	3	Sulphur	1 g
Filter paper strips	3	Dil. hydrochloric acid	10 ml
		Lead nitrate solution (1%)	2 ml

Procedure

Place the mercury in a mortar and add sulphur powder. There is no change in the colour of either, indicating that practically no reaction has occurred. Grind the two together for a while. A black powder (mercuric sulphide) is formed.

Take a little mercury, sulphur and the black powder formed in three separate test tubes. Add 2 ml of dilute hydrochloric acid to each. It will be found that while there is no change in the test tubes containing mercury and sulphur, there is a slight effervescence (due to the evolution of a gas smelling of rotten eggs) in the test tube containing the black powder. Hold a filter paper dipped in lead nitrate solution above the mouth of each test tube. The one held above the test tube containing the black powder turns black This indicates that the black substance is entirely different from mercury and sulphur and is formed by a reaction between these two.

Though contact between reactants is a primary requisite for reactions to occur, there are a number of reactions which do not take place by mere contact between reactants.

Supply of energy in various forms may be necessary to make such reactions occur. Demonstrations IV, V, and VI illustrate this point.

Demonstration IV Energy in the form of heat may be necessary to bring about a reaction

Time required : 30 minutes

Materials required

Hard glass test tube		Copper powder		2 g
Test tubes	3	Sulphur powder		3 g
Test tube holder		Dil. hydrochloric acid	(4M)	15 ml
Spirit lamp		Lead nitrate solution	(1%)	5 ml
Filter paper strips	4			

Procedure

Mix thoroughly about 2 g of the copper powder with about 3 g of powdered sulphur. Transfer the mixture to the hard glass test tube. Heat the mixture sufficiently for the reaction to set in (as indicated by the appearance of a glow). Stop heating. The reaction continues on its own and a black mass is formed. Cool thoroughly.

Take small quantities of powdered copper, sulphur and a mixture of the two separately in 3 test tubes. Add about 5 ml of dilute hydrochloric acid to each test tube and to the hard glass test tube containing the black substance. Warm all the four test tubes. There is no change in the test tubes containing copper, sulphur and the mixture of the two. But in the test tube containing the black substance there is an evolution of a gas smelling of rotten eggs Soak four filter papers in lead nitrate solution and hold each a little above the mouths of the four tubes. The filter paper held over the test tube containing the black substance will turn black while there will be no change in others. This indicates that the substance present in the hard glass test tube is different from either copper or sulphur. The conclusion, therefore, is that this new substance (copper sulphide) is formed as a result of the supply of energy to a mixture of powdered copper and sulphur.

Demonstration V Energy in the form of light brings about a chemical reaction

Time required : 20 minutes

Materials required

Beaker (150 ml)
Test tube
Funnel
Funnel stand
Glass rod
Filter paper
Dropper
Opaque object
(small iron strip)

Silver nitrate solution
(0.1 M) 10 ml
Potassium bromide solution
(0.1 M) 25 ml

Procedure

Fix a filter paper in a funnel. Soak it with potassium bromide solution. Let the excess of the solution drain. Add carefully, with a dropper, silver nitrate solution to the filter paper, in such a manner that a uniform layer of pale yellow silver bromide is formed on the filter paper (potassium bromide and silver nitrate react to give a precipitate of silver bromide). **Avoid undue exposure of silver bromide to light during its formation from potassium bromide and silver nitrate.** After the solution has completely drained out, remove the filter paper from the funnel. Cover a portion of the filter paper with a small opaque object. Expose the filter paper to sunlight for about 10 to 15 minutes. The exposed portion of the precipitate turns grey or black while that covered by the opaque object remains more or less unaffected (In the absence of sunlight the filter paper may be exposed to light from an electric lamp).

Demonstration VI Percussion brings about a reaction

Time required : 5 minutes

Materials required

Hammer	Potassium chlorate*	0 5 g
A piece of paper (5 cm. sq)	Sulphur*	0.2 g

* Use powdered samples.

Powder potassium chlorate in a clean mortar and transfer it to a clean, dry bottle. Clean the mortar with plenty of water and dry it. Make sure that it is free from any potassium chlorate. Then powder the sulphur.

Procedure

Mix carefully potassium chlorate and sulphur on a piece of paper. You will observe that no reaction takes place.

Wrap the mixture in the paper and place it on the floor. Strike it with a hammer. A loud explosion results and smoke issues.

The familiar example of Divali crackers may be brought to the attention of the pupils at this stage. These are tightly made packets containing a mixture of potassium chlorate and powdered sulphur along with small pebbles. When struck hard against a wall, or thrown on rocky ground, a loud explosion results

5.4 Factors which influence the rate of a reaction

No attempt should be made at any stage in this unit to define the rate of a reaction. Stress is to be laid on the commonsense understanding of the meaning of the term. For instance, a more brisk evolution of gas indicates a correspondingly faster reaction. It is presumed that this is self explanatory. Again quicker decolourisation of a coloured material may also be taken as indicative of a correspondingly faster reaction.

Experiment 19 Factors affecting the speed of a reaction

Time required 40 minutes

The experiments to be done here deal with the influence of factors like particle size, concentration, temperature and a catalyst on the rate of reaction.

A coverage of the relevant section in the text will serve as a postlab discussion.

Question and answer

What is the effect of

a. particle size
b. concentration
c. temperature and
d catalyst

on the rate of the reactions you have studied?

a. The rate of a reaction increases with decrease of particle size
b. The higher the concentration, the faster is the reaction.
c Rise of temperature increases the rate of reaction.
d. Catalyst increases the rate of the reaction.

Demonstration VII Particle size affects the rate of a reaction

Time required : 15 minutes

Materials required

Mortar and pestle	Potassium permanganate crystals	2 g
Dropper		
Watch glass	Glycerol (Glycerine)	5 ml

Procedure

Take about 1 g of coarse potassium permanganate on the watch glass. Add a few drops of glycerine. You will notice that hardly any reaction occurs.

Grind about a gram of the potassium permanganate to a fine powder. Transfer the powder to the watch glass and gather it into a small heap and flatten the top. Add a few drops of glycerine on the top of the heap The powdered mass kindles and glows.

Draw the attention of the pupils to the fact that the smaller the particle, the faster is the reaction.

Demonstration VIII Temperature influences the rate of a reaction

Time required : 30 minutes

Materials required

Beakers (150 ml)	3	Potassium permanganate solution	100 ml
Conical flasks (150 ml)	4	Oxalic acid solution	100 ml
Measuring cylinder (25 ml)		Dilute sulphuric acid	100 ml
Glass rods	2		

The following solutions are to be prepared in advance and kept ready for the demonstration.

Potassium permanganate solution

Dissolve about 0.3 g of crystalline potassium permanganate in 100 ml of distilled water in a beaker.

Oxalic acid solution

Dissolve about 0.7 g of crystalline oxalic acid in 100 ml of distilled water in a beaker.

Procedure

Label two conical flasks as A and B. Place 20 ml of the permanganate solution in each. Add 20 ml of dilute sulphuric acid to each conical flask.

Label two more conical flasks as C and D Take 20 ml of the oxalic acid solution in each. Heat conical flask C to about 65°C.

Add **simultaneously**, the hot oxalic acid solution C to the permanganate solution A and the cold oxalic acid solution D to the solution B. Stir the solutions.

It will be noticed that the decolourisation of potassium permanganate takes place more quickly with hot oxalic acid in flask A. This indicates that rise in temperature increases the speed of a reaction.

Types of chemical reactions

After all the experiments have been done by the pupils and the demonstrations shown, spend one period recalling the experiments done. Elicit information regarding different types of reactions from among the experiments done. Questions along the following lines may help.

1. What is the reaction between mangesium and oxygen to give magnesium oxide called ?

It is called combination, since magnesium and oxygen combine to form magnesium oxide.

2 Mercuric oxide, on heating, gives mercury and oxygen What type of reaction is this?

It is decomposition since mercuric oxide, on heating, breaks down to give simpler substances, mercury and oxygen.

3. Give another example of a decomposition reaction from the experiments you did?

Heating of copper carbonate, as it breaks into cupric oxide and carbon dioxide.

4. What type of reaction is the one between iron and copper sulphate solution ?

Displacement. In this reaction, iron displaces copper from copper sulphate solution giving copper and ferrous sulphate.

5 What is the reaction between sodium sulphate and barium chloride (in aqueous solution) known as?

Double decomposition. In this reaction there is a mutual exchange of ions leading to precipitation of barium sulphate.

5.4 Factors which influence the rate of a reaction

The study of the rates of chemical reactions and the factors that affect the rates is known as 'chemical kinetics'. The term 'rate of reaction' is used to describe the rate at which the reactants change to products. It may be defined as the change of concentration of the reactants per unit time. Changes in concentration can be determined directly or by measuring the change in some observable property of the reactants or products. Some of the properties made use of are, for instance, change in colour, evolution of gas, weight of a precipitate formed, change in refractive index etc.

The rate of a reaction varies from reaction to reaction. For example, when a solution of silver nitrate is added to a solution of sodium chloride a white precipitate of silver chloride is formed instantaneously. Such extremely fast reactions usually occur between ions and the rates of such reactions are difficult to measure. There are also extremely slow reactions, such as those occurring in geological processes. These may not reach completion even after thousands of years. A mixture of hydrogen and oxygen can be kept for years without any perceptible reaction taking place. A number of reactions which have been studied are those which are intermediate between these extremes.

A chemical reaction involves a rearrangement of atoms. This requires a weakening of the bonds holding the atoms together in the reactant molecules and the formation of new bonds in the products. The rate of a reaction depends, in large measure, upon the nature of the reactants But, there are also a number of other factors at our disposal which influence the rate, for instance, concentration, temperature, catalyst, etc.

Concentration

It is experimentally found that for homogeneous reactions increase of concentration of a reactant increases the reaction rate. (The reaction between nitric oxide and oxygen to form nitrogen dioxide is a homogeneous one, as the reactants and product are all in one phase). Occasionally, in a reaction involving two or more reactants, the rate is dependent on the concentration of some of the reactants only.

In heterogeneous reactions, the reactants are in more than one phase. In such cases, reactions occur only at the surface boundary between the reacting phases. The rate of reaction is proportional to the area of contact between the phases The burning of magnesium is a heterogeneous reaction involving magnesium (solid) and oxygen (gas). Rusting of iron is also a heterogeneous reaction involving iron (solid), water (liquid or vapour) and oxygen (in solution or in gaseous phase). It is slow when the surface of contact is small, as in the case of a bar of iron. If the bar is ground into powder, rusting is more rapid because of greater area of contact. Lumps of coal or zinc, for instance, are difficult to ignite in air, but when pulverised, they react vigorously and often explosively.

In order to react, molecules must come in close contact. Such contact takes place during collisions. In some of these collisions there is a rearrangement of atoms (and electrons), leading to the formation of products The **collision theory** serves as a useful model for explaining the effect of concentration upon reaction rate. In gaseous reactions, a simple way of increasing molecular collisions is to increase the pressure. Increase of pressure on a gas increases the density and so the concentration. This leads to an increase in the number of collisions per unit time. Hence rate of reaction increases.

A piece of wood burns more rapidly in pure oxygen than it does in air, in which oxygen is present only to the extent of about 20% by volume. Fanning a slow burning wood increases the concentration of oxygen next to the surface of the wood. Therefore it burns more rapidly.

Temperature

Temperature has a striking effect on the rates of reactions. Reactions which are negligibly slow at ordinary temperatures become extremely fast or occasionally explosive at higher temperatures. In most cases it is found that a 10°C rise in temperature approximately doubles the reaction rate.

The effect of temperature on reaction rate can be explained with the help of collision theory. But the theory requires modification even to qualitatively explain the temperature effect. The difficulties which the simple collision theory encounters can be visualised by considering a specific example – the decomposition of hydrogen iodide into hydrogen and iodine at some suitable temperature.

$$2HI \longrightarrow H_2 + I_2$$

Two molecules of HI should collide in order that the hydrogen and iodine atoms in the molecules come close enough to react and form H_2 and I_2. But the decomposition does not take place every time two HI molecules collide, for the observed rate of decomposition is much slower than it would be if every collision resulted in successful reaction. It is found that in a given sample of HI gas, the number of collisions taking place per unit time is far greater than the number of HI molecules which actually decompose in unit time This indicates that not all the collisions between reactant molecules bring about reaction. Evidently other factors are involved Among these 'other factors', the most important one is the **energy** of the reactant molecules. Some cosideration of the energy factors is therefore necessary to understand the effect of temperature upon reaction rate

The energy of a molecule

A molecule is a storehouse of energy. Part of it is the energy related to the various motions executed by the molecules and the remainder, the energy related to the forces which act between molecules and between the atoms in a molecule. The sum of all these energies is the total energy of the molecule. The energy possessed by one Avogadro number of molecules is called **molar heat content** of the substance. **It is the changes occuring in the heat contents of reactants and products that are manifested as** energy effects during chemical reaction.

The molecules in a gas are in a state of random motion, moving with different velocities. Therefore, the energy possessed by them varies from molecule to molecule. Only those collisions between reactant molecules which have energy over and above a certain minimum lead to the formation of products. Molecules having energy lower than this minimum rebound, without reacting, on collision. This minimum energy is called the **activation energy.**

The concept of activation energy is graphically illustrated in figure 5.2

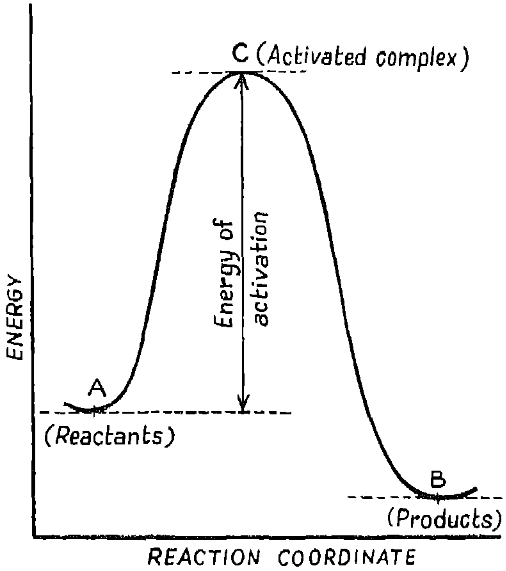


Figure 5 2 *Energy diagram for reaction :* $A \longrightarrow B$

In this diagram the average energy of the reactants and products in the reacting system is plotted against the reaction coordinate. The reaction coordinate tells us how far the reaction has proceeded from the initial to the final state A represents the average energy of the reactants and B that of the products. However, before the reactants can pass on to products they must cross an energy barrier. The reactant molecules having sufficient energy to react collide and travel up the energy barrier to C. C represents a transitory or intermediate stage formed by the collision of molecules with enough energy to react. This is known as the **'activated complex'. The activation energy is thus the energy required to convert the reactants into the activated complex.**

When the colliding molecules do not have enough energy to bring about a reaction, it simply means that they do not have sufficient energy to form the activated complex. But some fraction of the molecules have energy sufficient to form the activated complex If this fraction is small, reaction is slow, if it is large, the reaction is fast.

At any given temperature there is a constant fraction of molecules which possess the activation energy. A 10°C rise of temperature approximately doubles this fraction in most cases Hence the rate of reaction is also approximately doubled.

A small value of activation energy means a fast reaction, a high value means a correspondingly slower reaction

Catalyst

A catalyst usually provides an alternate path to the reaction, wherein the minimum energy required for reaction to occur is lowered. Hence the number of fruitful collisions between the reactants increases This results in a higher reaction rate.

Quiz

1. Explain how the particle size of reactants affects the rate of a reaction ?
2. In what form or forms is energy released when magnesium burns in air ?
3. Two compounds A and B are heated. A gives a gas which rekindles a glowing splinter. B gives a gas which turns lime water milky. What are the gases ? Which compounds mentioned in the text give these two gases on heating ?
4. Describe two combination reactions that are not mentioned in the text.
5. Why are the compounds of silver usually stored in dark coloured bottles ?
6. Give any two factors that would increase the rate of a reaction.
7. Zinc dust is added to two test tubes A and B, both of which contain an acid. Evolution of gas is more vigorous in A than in B. What conclusion do you draw from this ?

8. The action of heat on a piece of wood is called a chemical change while the action of heat on ice is called a physical change Why ?

9. Tick off (✓) the correct answer in the following :

a Increase in concentration of reactants
   i. increases
   ii. decreases
   iii. has no effect on

the rate of a reaction

b. The gas evolved on heating calcium carbonate is
   i. oxygen
   ii. hydrogen
   iii. carbon dioxide.

c. A reaction in which energy is absorbed is an
   i. exothermic reaction
   ii. endothermic reaction.

d. Decrease of temperature
   i. increases
   ii. decreases
   iii. has no effect on

the rate of a reaction.

6

Some Familiar Substances

INSTRUCTION MATERIAL

Introduction

This unit deals with some familiar substances like air, water, common salt, charcoal and sulphur. We shall also study the more important constituents of the above like oxygen, nitrogen, hydrogen and chlorine. All these substances occur widely in nature, possess wide range of properties and play an important role in our daily life

As seen earlier, air is a mixture mainly of oxygen and nitrogen. These constitute about one-fifth and four-fifths of the volume of air respectively. Besides these, air contains small amounts of carbon dioxide, water vapour and rare gases.

Of all the elements on the earth's surface, oxygen is the most abundant. It exists as molecular oxygen (O_2) in air; elsewhere it is found combined with other elements in the form of oxides. Free oxygen plays an important part in many processes, such as respiration, burning, the rusting of metals, decay in organic matter and the hardening of paints.

Nitrogen is extremely valuable to man and to all forms of life. Its presence in such large amounts dilutes the oxygen and slows down oxidation. Life in an atmosphere of pure oxygen would be quite different from what it is now. Combined nitrogen is always found as a constituent of the living matter in plants and animals.

The proportion of carbon dioxide in air is about 0.04 per cent. It is the product of respiration of all living beings. It is also obtained during the burning and decay of organic matter. But it is continuously used up in photosynthesis. Thus the concentration of carbon dioxide in the atmosphere remains fairly constant.

Next to air the most important substance necessary for the existence of life is water. It is the most abundant of all the chemical compounds. As seen earlier, it is a compound of hydrogen and oxygen.

In contrast to oxygen, very little free hydrogen is found in the atmosphere. Many important compounds such as acids and bases contain hydrogen as one of their constituents. Almost all organic compounds present in plant and animal tissues contain hydrogen in the combined form.Carbohydrates, fats and proteins, used as foods, contain hydrogen. Hydrogen is also found in combination with carbon in natural gas, petroleum and its products.

Common salt occurs in nature in the form of rock salt. Sea water also contains common salt to about 2.8 per cent. Common salt is a compound of sodium and chlorine. It is used in food preparations. Chlorine, one of the constituents of common salt, is used in the sterilization of drinking water. Chlorine is also used as a bleaching agent

Carbon is found in nature as the free element and in many compounds Carbon or carbon compounds, furnish us with practically all of our food, fuel and clothing. We depend upon carbon and its compounds to provide a large amount of energy used in our industrial activities such as extraction of metals from their ores. It is to be ranked along with oxygen and hydrogen as one of the three or four elements most important to man.

Sulphur is used in the manufacture of sulphuric acid, matches, explosives and medicines.

The chemical reactions of these substances are of fundamental importance and are necessary for understanding the chemistry of other elements. Hence, preparation, properties and uses of these substances are discussed in this unit.

Outline

1. Air (6 1)
2. Oxygen (6.2).
3. Nitrogen (6.3).
4. Carbon dioxide (6.4).
5. Water (6.5).
6. Hydrogen (6 6).
7. Common salt (6.7).
8. Chlorine (6.8).
9. Carbon (6.9).
10. Sulphur (6.10).

Concepts

1. Air is a mixture of nitrogen, oxygen, water-vapour, carbon dioxide and rare gases.
2. Catalysis is the process by which the rate of a chemical reaction is altered by the addition of a substance which itself does not undergo a permanent chemical change.
3. Oxidation is a process by which substances combine with oxygen. Substances like potassium chlorate which give oxygen for oxidation are called oxidising agents.
4. Reduction is removal of oxygen from a compound. A substance which removes oxygen is called a reducing agent.
5. Certain elements like carbon and sulphur can exist in different forms with different physical properties like crystalline shape, colour, density, solubility, etc

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	20	—	I	6.1
2	—	Discussion	II	6.2
3	21	—	III	6.3
4	22	—	IV	6.4
5	23	—	V	6.5 and 6.6
6	24	—	VI	6.7 and 6.8
7	25	Discussion	VII	6.9
8	26	Discussion	VIII	6.9 and 6.10
9	—	—	IX	6.10
10	—	—	—	Review and quiz

Development

6.1 Air

The first lesson consists of demonstration I followed by discussion and laboratory experiment 20.

Before proceeding to demonstration I, the bell jar experiment demonstrated in unit 1 may be recalled and the pupils reminded that oxygen and nitrogen are the two major components of air. The pupils may then be asked whether any other gases are present in air. Recall from their experiences in the study of biology, what things are necessary for the preparation of starch by plants.

After initiating the above discussion, proceed to demonstration I (The aim of the demonstration I may not be revealed to the pupils in the beginning).

Demonstration I Composition of air

Time required : 50 minutes

I A To show that air contains carbon dioxide

Materials required

Beaker (100 ml)	Lime water 20 ml
Cycle pump	

Procedure

Take 20 ml of lime water in a beaker. Using a cycle pump, blow air into the lime water for about 15 minutes.

Lime water turns milky indicating the presence of carbon dioxide in air.

Now the pupils know that air contains carbon dioxide, in addition to oxygen and nitrogen. Ask the pupils whether any other substances are also present in air. You may recall the rusting of iron when exposed to air. Is rusting only due to oxygen or any other substance present in the air ?

This is the proper time for demonstration I B to be taken up.

I B To show that air contains water vapour

Materials required

Beaker (100 ml)
Watch glass
Crushed ice

Procedure

Wipe the outer walls of the beaker. Fill it three-fourths with crushed ice and cover it with a watch glass. Set aside the beaker for about 15 minutes.

Water droplets are found on the outer walls of the beaker indicating the presence of water vapour in air.

The formation of rain, fog, mist and dew may be cited in showing the presence of water vapour in the atmosphere. Tell the pupils that William Ramsay and others showed that air also contains a small proportion of rare gases · helium, argon, neon, krypton and xenon.

6.2 Oxygen

Experiment 20 Preparation and properties of oxygen

Time required : 20 minutes

Before asking the pupils to carry out experiment 20, it will be necessary to explain to them the operations involved in the preparation and collection of gases, such as clamping the tube and collecting the gas in water filled test tubes. The experiment may be carried out by the pupils in pairs. All the materials required must be kept before the laboratory session starts. The pupils must be allowed to assemble the apparatus themselves with the help of instructions and figures given in the laboratory manual.

Question and answer

In this experiment oxygen is collected over water. From this what can you infer about the solubility of oxygen in water?

From the fact that oxygen is collected over water, it can be said, that oxygen is not appreciably soluble in water

The second lesson consists of discussion on experiment 20 followed by demonstration II. The discussion on experiment 20 may be started by asking one of the pupils how he prepared oxygen. Pose the question–where from has the oxygen come? Obviously from potassium permanganate.

What type of reaction has taken place resulting in the evolution of oxygen?

Their knowledge of chemical reactions in unit 5 must be helpful in answering the above question. Elicit from the pupils that potassium permanganate decomposes on heating, giving out oxygen. Let them also realise that since potassium permanganate yields oxygen, it must contain oxygen. Certain other substances containing oxygen can also be decomposed by heating to yield oxygen. Tell the pupils that one such substance is potassium chlorate.

At this stage, demonstrate that the evolution of oxygen is very slow when potassium chlorate alone is heated (test with a glowing splinter). Now mix a little manganese dioxide (free from carbon) with potassium chlorate and heat the mixture. Let the pupils observe that the oxygen is evolved more readily (rekindling of the glowing splinter).

Also tell the pupils that the oxygen liberated does not come from the manganese dioxide as it does not decompose at this temperature. Furthermore, after the reaction is completed, manganese dioxide can be separated from potassium chlorate (in the residue) by taking advantage of the fact that the latter is soluble in water while the former is not The quantity of manganese dioxide recovered in this way is equal to the amount originally added. You may conclude, therefore, that manganese dioxide only increases the speed of the decomposition of potassium chlorate and is not consumed during the reaction. A substance which influences the speed of the reaction, without undergoing any permanent chemical change itself, is called a catalyst (unit 5).

Demonstration II Preparation and properties of oxygen

Time required : 50 minutes

Materials required

Iron stand with adjustable clamp	
Hard glass test tube	
Rubber stopper carrying a delivery tube	
Glass trough	
Evaporating dish	
Gas jars with covers	5
Beehive shelf	
Reagent bottle with cork (4 oz)	
Spirit lamp	
Box of matches	
Deflagrating spoon	
Tongs	
Sand paper	

Potassium chlorate	10 g
Manganese dioxide	3 g
Splinters	2
Charcoal pieces	
Sulphur	0.5 g
Lime water	5 ml
Magnesium ribbon	5 cm
Copper wire	5 cm

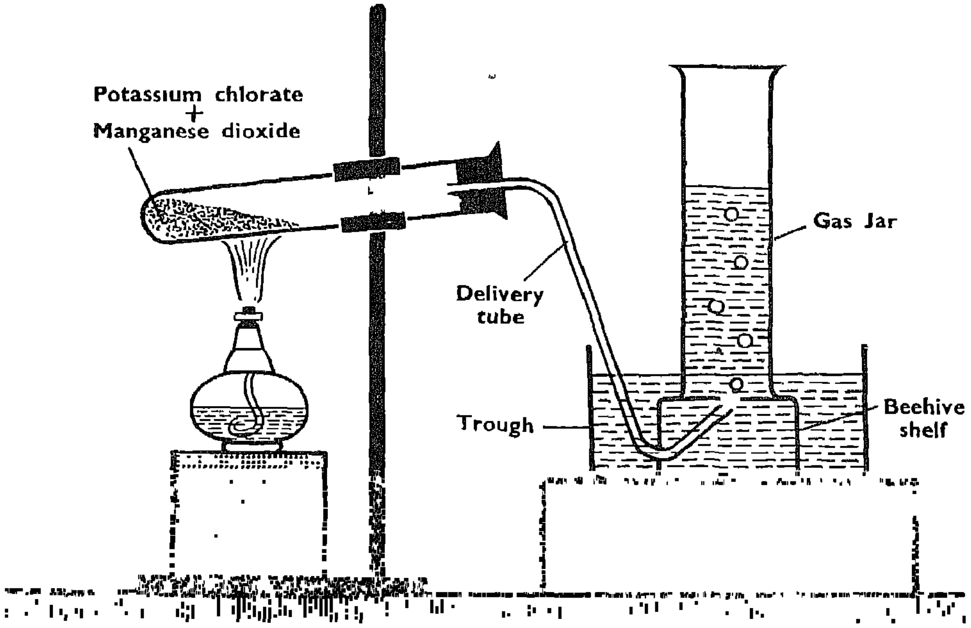


Figure 6.1 *Preparation of oxygen*

II A Preparation of oxygen

Oxygen is usually prepared by heating potassium chlorate in the presence of manganese dioxide which acts as a catalyst

Procedure

Ignite 3 g of manganese dioxide in an evaporating dish to burn off any carbon present in it as an impurity. **If carbon is not removed, it may cause explosion when it is heated with potassium chlorate.**

Mix thoroughly 10 g of potassium chlorate with the pre-heated manganese dioxide in a mortar. Take this mixture in a hard-glass tube. Fit it with one holed rubber stopper carrying a delivery tube. Spread out the mixture as shown in figure 6.1.

Gently heat the thin layer of the mixture first, so that there is a slow and steady evolution of oxygen. Let the first few bubbles of air escape.

Fill a gas jar completely with water and cover it with a lid, Make sure that the lid fits firmly against the mouth of the jar. If there are any air bubbles, add water and cover again. Invert the gas jar and keep it under water in the trough filled upto two-thirds. Remove the lid from the mouth of the gas jar under water and keep it over a beehive shelf placed at one end of the trough Then insert the end of the delivery tube into the beehive shelf. Now the gas will collect in the gas jar by the displacement of water. (A gas that does not dissolve in water is usually collected by downward displacement of water). As soon as the jar is completely filled with oxygen, close the mouth of the jar with the lid. Then remove the jar from the water and keep it mouth upwards (as oxygen is slightly heavier than air). In a similar manner, collect oxygen in four more gas jars. (It is convenient to have 3 or 4 gas jars filled with water and kept in an inverted position under water in the trough. This will enable collection of oxygen without any wastage. To avoid overflow remove water from the trough as oxygen is being collected).

To test the solubility of gas in water, collect the gas in a 4 oz. reagent bottle and cork it.

When the gas jars are filled with oxygen, remove the delivery tube from the water trough and stop heating.

$$2KClO_3 \xrightarrow[MnO_2]{heat} 2\,KCl + 3O_2\uparrow$$

II B Properties of oxygen

To study properties of oxygen and show that oxygen is reactive.

General instructions

1 Test the solubility of the gas as follows :

Place the mouth of the corked gas bottle under water in a trough, remove the cork, shake the bottle to ensure mixing of gas with water and wait for several minutes to allow time for solution. Observe the height of the water in the bottle. Height of water indicates the solubility. If the gas is appreciably soluble in water, the height of the water level in the bottle is more, if moderately soluble, the height is less and if slightly soluble the height is least.

2. Keep the gas jars of oxygen mouth upwards Remove the cover of the gas jar and do the following reactions.

3. For each test, use a dry and clean deflagrating spoon.

	Test	Observation
1.	State and colour	Colourless gas
2	Odour	Odourless
3.	Solubility in water	Slightly soluble (as the height of the water level in the bottle is very small)
4	**Testing with splinter**	
a.	Introduce a burning splinter into the jar of oxygen.	Burning splinter burns brightly and the oxygen gas does not burn.
b	Introduce a glowing splinter into the same jar of oxygen.	Glowing splinter is rekindled and burns brightly This indicates that oxygen helps in the burning of the splinter.
5.	**Burning of charcoal** Hold a piece of charcoal with a pair of tongs and heat it in a flame until it glows. Place the glowing charcoal in a deflagrating spoon and introduce it into the jar of oxygen.	Glowing charcoal is rekindled and burns with an orange flame forming a colourless gas
	After sometime remove the deflagrating spoon from the gas jar. Add 5 ml of lime water to the gas jar and shake.	Lime water turns milky showing the gas formed by the burning of charcoal to be carbon dioxide
6.	**Burning of sulphur**	
	Take about 0.5 g of sulphur in a deflagrating spoon. Heat the spoon in the flame until the sulphur melts and just starts burning. Then introduce the deflagrating spoon into the jar of oxygen	Burning sulphur continues to burn with a blue flame and forms a gas having suffocating odour (sulphur dioxide).

7. **Burning of magnesium**

Hold a clean piece of magnesium ribbon (5 cm) with a pair of tongs and heat till it starts burning. Introduce the burning magnesium ribbon into the jar of oxygen.	Burning magnesium continues to burn with a dazzling light and forms white ash of magnesium oxide.

8. **Burning of copper**

Hold a piece of copper wire (5 cm) with a pair of tongs and heat it in a flame until it is red hot Introduce the red hot copper wire into the jar of oxygen.	Red hot copper wire burns and forms a black substance (cupric oxide).

Discussion

Summarise, in a tabular form, the results of the reactions carried out with oxygen.

Point out that most of the substances such as charcoal, sulphur, magnesium and copper burn in oxygen forming the corresponding oxides.

Now explain the term oxidation. The combination of a substance with oxygen is called oxidation. The substances such as potassium permanganate and potassium chlorate which supply oxygen for oxidation are called oxidising agents.

6.3 Nitrogen

The third lesson consists of experiment 21 followed by demonstration III.

Experiment 21 Preparation and properties of nitrogen

Time required : 20 minutes

Question and answer

Is nitrogen soluble in water ?

Nitrogen is slightly soluble in water.

Demonstration III Preparation and properties of nitrogen

Time required : 60 minutes

Materials required

Iron stand with clamp	Sodium nitrite	10 g
Round bottomed flask (250 ml)	Ammonium chloride	10 g

One holed rubber stopper	
Delivery tube	
Gas jars with covers	5
Trough	
Beehive shelf	
Deflagrating spoon	
Tongs	
Measuring cylinder (25 ml)	
Spirit lamp	
Box of matches	
Charcoal piece	
Sulphur	0.5 g
Magnesium ribbon	5 cm
Copper wire	5 cm
Splinter	

III A Preparation of nitrogen

Nitrogen is obtained by the decomposition of ammonium nitrite, NH_4NO_2. Since ammonium nitrite is very unstable, some danger is involved in heating the pure and dry substance In order to avoid this danger, the substance is usually prepared in situ from a mixture containing sodium nitrite and ammonium chloride in solution.

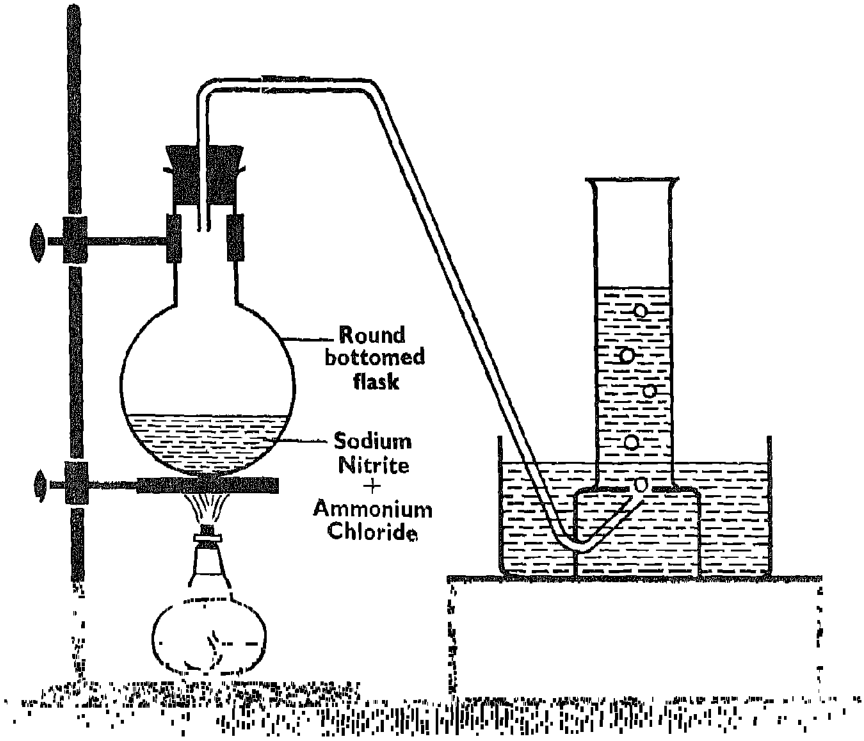


Figure 6.2 *Preparation of nitrogen*

Procedure

Take about 10 g of sodium nitrite ($NaNO_2$) and 10 g of ammonium chloride (NH_4Cl) in a round-bottomed flask. Add 50 ml of water. Fit the flask with a cork carrying a delivery tube as shown in figure 6.2. Heat the flask for about 5 minutes or until there is an evolution of gas Stop heating and collect the nitrogen in five gas jars after allowing the first few bubbles of air to escape (See Demonstration I for details). To test the solubility of gas in water, collect the gas in a reagent bottle (4 oz) and cork it.

Caution . Strong heating must be avoided as it might result in an explosion.

$$NH_4Cl + NaNO_2 \longrightarrow NH_4NO_2 + NaCl$$

$$NH_4NO_2 \longrightarrow 2H_2O + N_2\uparrow$$

III B Properties of nitrogen

To study the properties of nitrogen and show that it is not a very reactive gas.

Note the colour and odour. Test its solubility in water, and introduce a burning splinter, charcoal, sulphur and magnesium and red hot copper wire into the jars containing nitrogen. Record the observations in a tabular form on the board

Discussion

Point out that most of the substances such as charcoal, sulphur and copper do not burn in nitrogen. However, burning magnesium continues to burn in nitrogen and forms magnesium nitride.

6.4 Carbon dioxide

The fourth lesson consists of experiment 22 followed by demonstration IV.

Experiment 22 Preparation and properties of carbon dioxide

Time required : 20 minutes

Question and answer

Is carbon dioxide heavier or lighter than air ?

Carbon dioxide is heavier than air.

Demonstration IV Preparation and properties of carbon dioxide

Time required : 60 minutes

Materials required

Round bottomed flask (250 ml)		Marble chips	25 g
Thistle funnel		Conc. hydrochloric acid	10 ml
Two holed rubber stopper		Sulphur	1 g
Delivery tube		Magnesium ribbon	5 cm
Tongs		Lime water	5 ml
Gas jars with covers	5	Splinters	2
Reagent bottle (4 oz)			
Candle			
Spirit lamp			

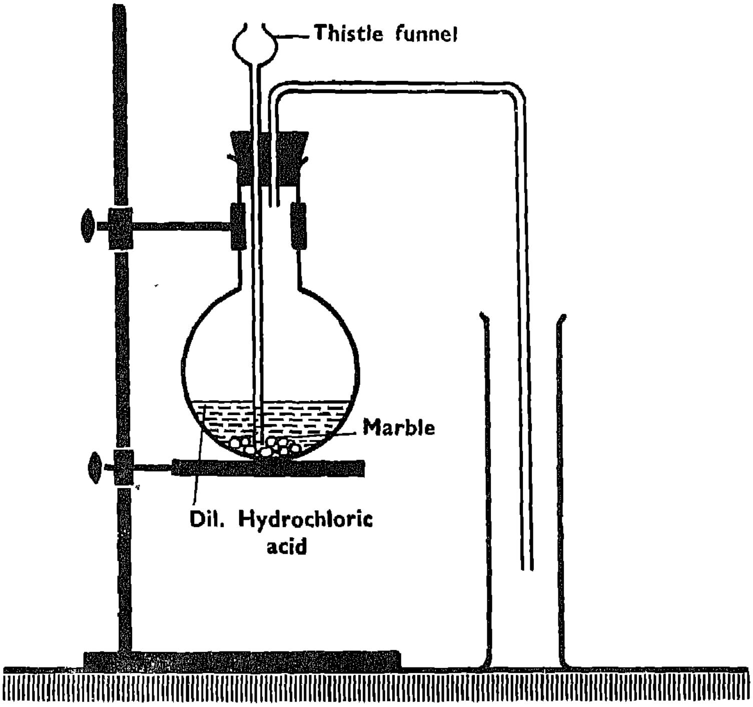


Figure 6 3 *Preparation of carbon dioxide*

IV A Preparation of carbon dioxide

Carbon dioxide is prepared by the action of dilute hydrochloric acid on marble (calcium carbonate).

Procedure

Place about 25 g of marble chips, $CaCO_3$, in the flask. Fit the flask with two holed rubber stopper carrying a thistle funnel and a delivery tube as shown in figure 6.3.

Pour water down the funnel to cover the marble chips and the lower end of the thistle funnel. The thistle funnel serves two functions: acids or other liquids may be added through it; it acts as a safety valve when the amount of gas produced is more than that can escape through the delivery tube. (A dropping funnel can be used instead of a thistle funnel and the acid can then be added drop by drop to produce the gas steadily). Add 10 ml of concentrated hydrochloric acid until carbon dioxide is formed at a moderate rate.

Wait till the gas displaces all the air from the apparatus. Then collect 5 gas jars of carbon dioxide by the upward displacement of air. Test with burning splinter to ascertain whether the jar is filled with carbon dioxide. Collect the gas in the reagent bottle to test its solubility and cork it.

Cover the gas jars firmly with glass squares and keep them mouth upwards.

$$CaCO_3 + 2HCl \longrightarrow CO_2\uparrow + CaCl_2 + H_2O$$

IV B Properties of carbon dioxide

Note the colour and odour, and test its solubility in water.

Show that carbon dioxide is denser than air by pouring it into an empty gas jar containing a burning candle.

Introduce a burning splinter, sulphur, and magnesium into the jar of carbon dioxide

Test the gas with lime water. Add 5 ml of clear lime water to a jar of carbon dioxide. Close it with a lid and shake well.

Discussion

Carbon dioxide, like nitrogen, is not highly reactive. It is not combustible. It puts out a burning splinter and sulphur. However, burning magnesium is hot enough to decompose carbon dioxide. Thus a piece of burning magnesium ribbon continues to burn in a jar of the gas. The magnesium unites vigorously with the oxygen set free by the decomposition and carbon is produced, as shown by a coating of soot on the inside of the jar.

$$2\,Mg + CO_2 \longrightarrow 2\,MgO + C$$

Carbon dioxide turns lime water milky due to the precipitation of calcium carbonate. This reaction is used as a test for carbon dioxide. If excess carbon dioxide gas is bubbled through the solution, the precipitate disappears due to the formation of soluble calcium hydrogen carbonate, $CaHCO_3$.

6.5 Water

The fifth lesson deals with the importance of water and one of its constituents, hydrogen. You may recall the electrolysis of water (unit 1) and tell the pupils that water consists of hydrogen and oxygen. The pupils are already familiar with the preparation and properties of oxygen. So they may be asked to take up experiment 23. This is followed by demonstration V with discussion.

6.6 Hydrogen

Experiment 23 Preparation and properties of hydrogen

Time required : 20 minutes

Question and answer

Is hydrogen heavier or lighter than air ?
Hydrogen is lighter than air.

Demonstration V Preparation and properties of hydrogen

Time required : 50 minutes

Materials required

Iron stand with clamp		Zinc granules	8 g
Round bottomed flask (250 ml)		Copper sulphate solution (1%)	1 ml
Thistle funnel		Conc. sulphuric acid	4 ml
Two holed rubber stopper		Splinters	2
Delivery tube			
Trough			
Beehive shelf			
Gas jars with covers	2		
Measuring cylinder (10 ml)			

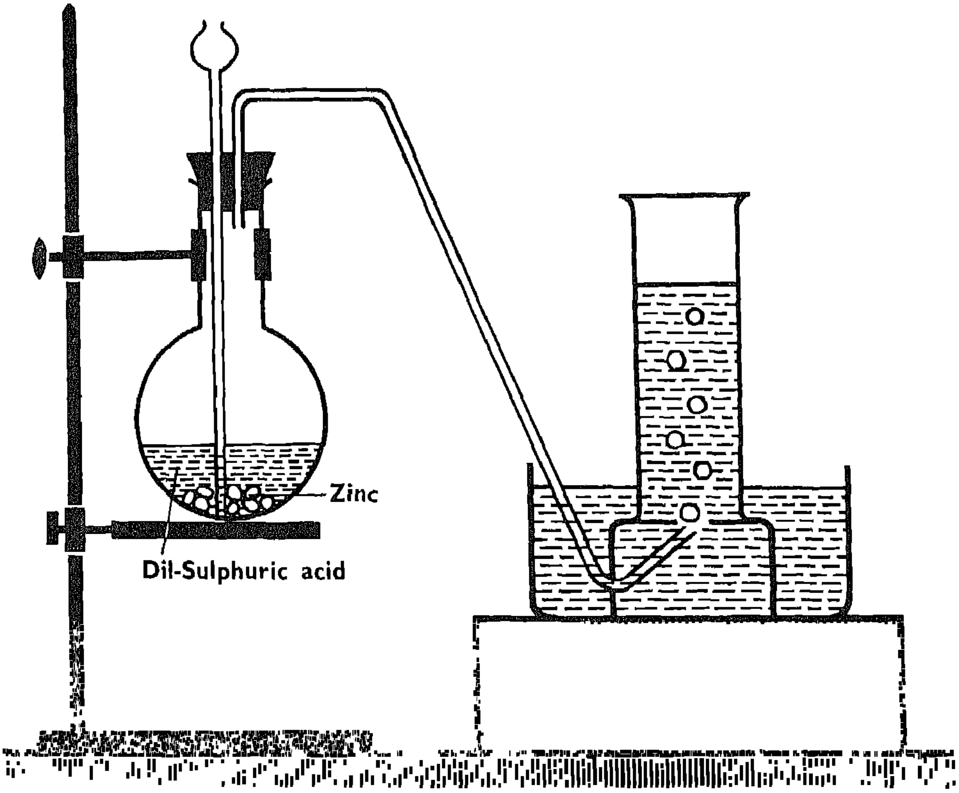


Figure 6.4 *Preparation of hydrogen*

V A Preparation of hydrogen

Hydrogen is prepared by the displacement of hydrogen present in the acid by metals like zinc The gaseous hydrogen so obtained is collected by the displacement of water.

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow$$

Procedure

Take about 8 g of granulated zinc in a round bottomed flask. Add enough water to cover the zinc. Fit the flask with a two holed rubber stopper carrying a thistle funnel and a delivery tube, as shown in figure. 6.4 Make sure that the end of the thistle funnel reaches the bottom of the flask.

Now add, through the thistle funnel, about 4 ml of conc. sulphuric acid and a few drops of copper sulphate solution (the latter acts as a catalyst in this reaction). Let the pupils observe the formation of the gas bubbles on the surface of the zinc granules They should also note that the flask becomes warmer as the reaction proceeds

After allowing the air bubbles to escape, collect the hydrogen in two gas jars by the displacement of water. Close the gas jars with covers and keep them mouth downwards (as hydrogen is lighter than air). Collect the gas in a reagent bottle to test its solubility and cork it.

1. **As a mixture of air and hydrogen has a tendency to explode when ignited, it is essential that all the air must be allowed to escape before collecting the hydrogen.**
2. **The first jar of hydrogen collected might still contain some air. It is therefore suggested that it may be used only to study the physical properties.**
3. **For the same reason, do not have any flame near the apparatus during the preparation of hydrogen.**

Note :

1. If the evolution of the gas slows down while collecting the gas, add a few ml of concentrated sulphuric acid to keep the reaction going
2. Note that all metals and acids cannot be employed to generate hydrogen. Metals like gold, silver and copper are not active enough to replace the hydrogen from acids. Nitric acid and hot concentrated sulphuric acid are strong oxidising agents; hence, in the reaction of these acids with metals, any hydrogen formed is immediately oxidised to water.

V B Properties of hydrogen

Carry out the following tests and record the observations in a tabular form.

1. Colour and odour.
2. Test solubility in water.
3. *Density relative to air*
   Keep a jar of hydrogen mouth upwards. Remove the cover. Wait for one minute. Then introduce a burning splinter.
4. *Testing with a splinter*
   Insert a burning splinter into a jar containing hydrogen (hold the jar of hydrogen mouth downwards).

V C Hydrogen as a reducing agent

Materials required

- Iron stand with clamps
- Tripod
- Hydrogen generator (Flask fitted with delivery tube and thistle funnel)
- Hard glass tube
- Test tubes 2
- Spirit lamp
- Box of matches

Zinc granules	30 g
Conc sulphuric acid	10 ml
Conc. nitric acid	5 ml
Cupric oxide	0.5 g
Copper sulphate	1 g
Copper turnings	0.5 g

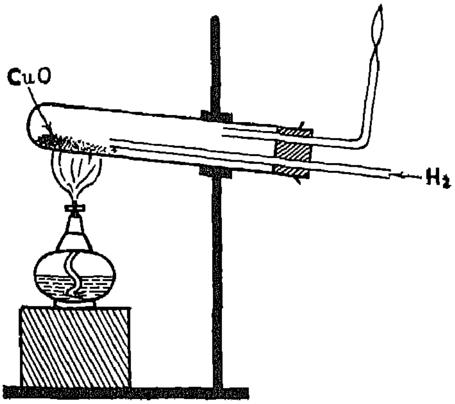


Figure 6.5 *Reduction of cupric oxide*

Procedure

Place about 0.5 g of cupric oxide in a hard glass tube and set up the apparatus as shown in figure 6.5

Pass a slow stream of hydrogen from the generator (refer the demonstration V A for details) into the hard glass tube containing cupric oxide for about two minutes to ensure complete displacement of air. Otherwise a mixture of hydrogen and oxygen (from air) will explode when heated

Now heat the cupric oxide for about 10 minutes, passing hydrogen over it The cupric oxide begins to glow. After 10 minutes stop heating. Continue to pass hydrogen until the glass tube cools to room temperature. The presence of a reddish-brown substance is noticed in the centre of the glass tube. Some liquid has condensed as droplets at the other end of the tube. Stop heating and continue passing hydrogen.

Note :

The hot metallic copper formed after reduction, might get oxidised to cupric oxide, if allowed to cool in air. To prevent this, the hot copper is allowed to cool in an atmosphere of hydrogen.

Elicit that the droplets formed are water ; sprinkle a small amount of white anhydrous copper sulphate over the droplets The white anhydrous copper sulphate turns blue. Show that anhydrous copper sulphate turns blue when water is added.

(Prepare anhydrous copper sulphate by heating 1 g of blue crystalline copper sulphate till it becomes white).

Elicit that the reddish-brown substance is copper. Remove carefully the reddish-brown substance from the glass tube. Take it in a test tube and add 2 ml of concentrated nitric acid. Brown fumes are evolved and the solution in the test tube becomes blue. Compare this with the action of nitric acid separately on copper and cupric oxide.

Add 1 ml of concentrated nitric acid to about 0.5 g of copper turnings taken in a test tube. Brown fumes are evolved and a blue solution is obtained.

Add 1 ml of concentrated nitric acid to a small amount of cupric oxide taken in a test tube No reaction takes place.

Discussion

In this demonstration, you find that on passing hydrogen over cupric oxide, the latter is converted into copper, and water is formed. From cupric oxide oxogen is removed by hydrogen.

$$CuO + H_2 \longrightarrow Cu + H_2O$$

Removal of oxygen from a compound is called reduction. Hydrogen which reduces cupric oxide to copper is called a reducing agent. During this process hydrogen gets oxidised to water.

6.7 Common salt

The sixth lesson deals with the importance of common salt and one of its constituents, chlorine. You may recall the formation of sodium chloride from sodium and chlorine (unit 1). The pupils may be asked to take up experiment 24 which will be followed by demonstration VI

6.8 Chlorine

Experiment 24 Preparation and properties of chlorine

Time required : 20 minutes

Question and answer

Is chlorine heavier or lighter than air ?

Chlorine is heavier than air.

Demonstration VI Preparation and properties of chlorine

Time required : 60 minutes

Materials required

- Iron stand with clamp
- Round bottomed flask (250 ml)
- Thistle funnel
- Two holed cork
- Delivery tube
- Beehive shelf
- Deflagrating spoon
- Test tube
- Trough
- Gas jars with covers 6
- Card board with a whole in the centre
- Measuring cylinder (25 ml)
- Candle
- Spirit lamp

Manganese dioxide	8 g
Conc. hydrochloric acid	40 ml
A jar of hydrogen	
Sodium	0.5 g
Magnesium ribbon	5 cm
Iron filings	0.5 g
Blue litmus paper	
Petal of a red flower	

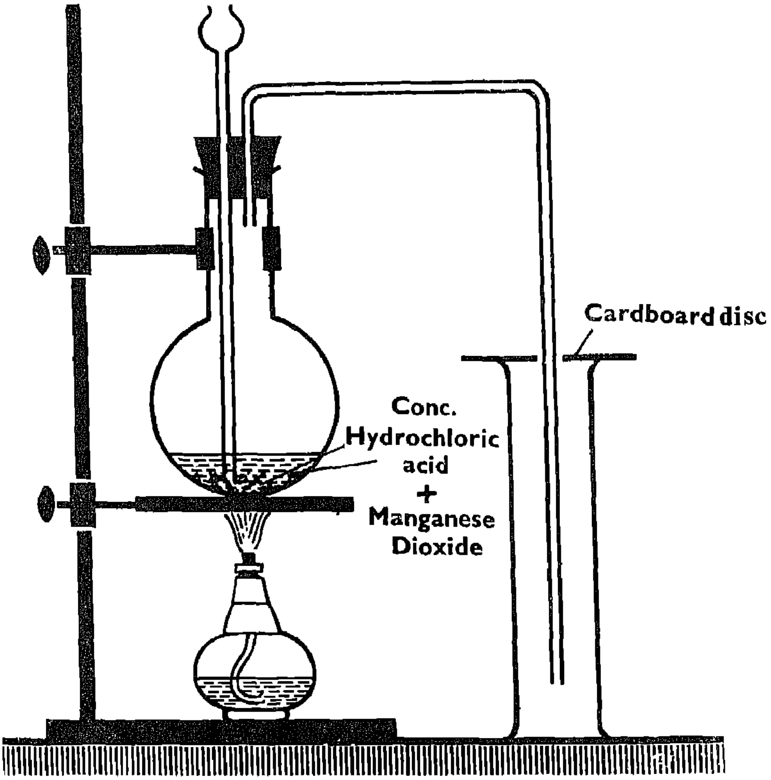


Figure 6.6 *Preparation of chlorine*

VI A Preparation of chlorine

Chlorine is prepared by the oxidation of hydrochloric acid using manganese dioxide as an oxidising agent.

Procedure

Place 8 g of manganese dioxide in the round bottomed flask. Fit the flask with the two holed cork carrying a thistle funnel and a delivery tube as shown in figure 6 6. Add 40 ml of concentrated hydrochloric acid through the funnel. Heat the flask for a few minutes Collect the gas evolved in a gas jar by the upward displacement of air. While the gas is being collected, the gas jar is kept covered with a piece of card board (the delivery tube can pass through the hole in the card board).

Hold a piece of moistened blue litmus paper just above the mouth of the gas jar. When the litmus paper is bleached, the gas jar will be full of chlorine, and the jar may be removed. Close the gas jar with the lid and keep it mouth upwards.

In a similar manner, collect chlorine in six gas jars.

$$MnO_2 + 4\,HCl \longrightarrow MnCl_2 + 2\,H_2O + Cl_2\uparrow$$

Note :

When the collection of chlorine is complete, remove the flame and put the end of the delivery tube in a beaker filled with water (or preferably 100 ml of 5% sodium thiosulphate solution) to absorb the excess of chlorine and thus prevent its escape into the room.

Caution :

As chlorine is a poisonous gas it should not be inhaled. So, while collecting it avoid overflow from the gas jars.

VI B Properties of chlorine

Carry out the following tests and record the observations in a tabular form.

1. Colour and odour.
2. Test the solubility in water.
3. *Reaction with hydrogen in diffuse light*

   Invert a covered jar of chlorine over a covered jar of hydrogen Remove the covers. Wait for 2 minutes.

Caution :

Direct sunlight may cause an explosion. So, keep the jars in diffuse light only.

Separate the jars and keep them mouth upwards. Bring a glass rod dipped in concentrated ammonium hydroxide solution near the mouth of each jar.

4. Test with sodium.
5. Test with burning magnesium.
6. *Reaction with iron filings*

   Heat the iron filings in the test tube and drop them into the jar of chlorine.
7. *Reaction with burning candle*

   Cut out a small piece of candle Light it, place it in a deflagrating spoon and lower it into a jar of chlorine.
8. *Bleaching of coloured material*

   Introduce a blue litmus paper or a petal of a red flower (free from moisture) into a jar of chlorine. Add a few drops of water into the jar and shake

Discussion

Chlorine reacts with hydrogen forming hydrogen chloride gas. Sodium, magnesium and iron filings burn in chlorine and form their respective chlorides. Chlorine bleaches the colouring matter in the presence of moisture and acts as a bleaching agent.

6.9 Carbon

The seventh lesson consists of experiment 25 followed by demonstration VII with discussion.

Experiment 25 Properties of carbon

Time required . 20 minutes

Question and answer

What gas is formed when cupric oxide is heated with charcoal?

Carbon dioxide is formed.

Discussion

You may recall here that charcoal when burnt in oxygen forms carbon dioxide. In the reaction 3, carbon dioxide is formed when carbon is heated with cupric oxide Cupric oxide supplies oxygen and acts as an oxidising agent.

Demonstration VII Properties of carbon

Time required : 50 minutes

VII A Carbon as an adsorbing agent

Materials required

Tripod		Gas jar containing nitrogen dioxide*	
Wire gauze		Powdered charcoal	2 g
Tongs		Blue litmus solution	50 ml
Beakers (150 ml)	2	Animal charcoal	1 g
Funnel			
Funnel stand			
Test tube			
Spirit lamp			
Box of matches			

***Preparation and collection of nitrogen dioxide**

Add 5 ml of concentrated nitric acid to about 2 g of copper turnings taken in a test tube. Immediately fit the test tube with one holed cork carrying a delivery tube. Keep the end of the delivery tube in a gas jar and collect brown coloured nitrogen dioxide gas (obtained by the action of concentrated nitric acid on copper turnings). When the gas jar is full of brown coloured gas, remove the end of the delivery tube from the gas jar and cover the gas jar with a plate and use this for demonstration VII A.

Adsorption of nitrogen dioxide gas

Procedure

Remove the plate of the gas jar containing reddish-brown nitrogen dioxide.

Add 2 g of powdered charcoal. Close the jar with the plate and shake the jar well (holding tightly the plate). Reddish-brown colour disappears indicating that powdered charcoal has adsorbed nitrogen dioxide

Now transfer the charcoal from the gas jar into a test tube and heat gently Reddish-brown fumes (of nitrogen dioxide) are given out.

VII B Decolourisation of litmus solution

Procedure

Take about 50 ml of water in a beaker. Add about 5 ml of blue litmus solution to it. Then add about 5 g of animal charcoal to the solution and boil for five minutes. Filter the mixture and collect the filtrate in a clean beaker. Filtrate is colourless showing that animal charcoal has adsorbed the blue colour of the solution.

Note :

Do a parallel test without adding animal charcoal and show that the colour of the solution remains unchanged.

Discussion

Charcoal powder has large surface area per gram, and so it is able to hold up nitrogen dioxide gas and other colouring substances like litmus . This process is called adsorption.

The eighth lesson consists of demonstration VIII with discussion followed by experiment 26.

Demonstration VIII Carbon as a reducing agent

Time required 40 minutes

Materials required

- Beaker (150 ml)
- Hard glass test tube
- Test tube
- Test tube holder
- Glass rod
- Box of matches
- Spirit lamp
- Filter paper

Cupric oxide	1 g
Powdered charcoal	1 g
Conc. nitric acid	5 ml

Procedure

Take about 1 g of cupric oxide and 1 g of powdered charcoal in a mortar and grind them thoroughly with a pestle. Place the mixture in a clean and dry hard glass test tube. Hold it with a test tube holder. Heat the test tube strongly for about 5 minutes. Then introduce a burning splinter into the test tube. The burning splinter is put out.

Continue heating for another 10 minutes Cool the test tube. Transfer the contents, of the test tube into a beaker containing about 75 ml of water, Stir well the contents of the beaker with a glass rod and set it aside for sometime. The presence of a reddish-brown substance may be noticed in the beaker. Decant carefully the liquid portion containing the floating charcoal powder. Wash the residue at the bottom of the beaker two or three times, by decantation, to remove the charcoal powder and any unreacted cupric oxide. Collect the reddish-brown particles on a filter paper and dry them (using bits of filter paper). Transfer the reddish-brown particles into a test tube and add 1 ml of concentrated nitric acid. Brown fumes are evolved and the solution in the test tube turns greenish blue.

This indicates that the reddish-brown particles are of metallic copper.

Discussion

The carbon (charcoal) reacts with the oxygen of cupric oxide when heated. In this process, carbon acts as a reducing agent and reduces cupric oxide (by removing oxygen) to metallic copper. Carbon is oxidised to carbon dioxide. Thus metallic copper and carbon dioxide are formed during the reduction of cupric oxide by charcoal.

$$2\,CuO + C \longrightarrow 2\,Cu + CO_2\uparrow$$

The other two forms of carbon — diamond and graphite and their properties and uses may be described to pupils after the above discussion.

6.10 Sulphur

Experiment 26 Properties of sulphur

Time required 30 minutes

Discussion

You may recall the test 2 in experiment 26 and inform the pupils that the effect of heat on sulphur is an interesting phenomenon.

On heating, sulphur melts to an amber coloured mobile liquid. As the temperature is raised, it becomes much darker and viscous.

It becomes so viscous that the test tube may be inverted without spilling the sulphur. On further heating, the sulphur gradually becomes mobile again and results in a dark reddish-brown liquid.

The ninth lesson consists of demonstration IX, on forms of sulphur with discussion.

Demonstration IX Forms of sulphur

Time required : 80 minutes

IX A Preparation of monoclinic sulphur

Materials required

Evaporating dish
Glass rod
Pen-knife
Tongs

Sulphur	15 g

Procedure

Place powdered sulphur in an evaporating dish. Heat it and stir, gradually adding more sulphur until the dish is almost brim-full of molten sulphur Use a small flame for the heating (or the sulphur may begin to burn). Then allow the sulphur to cool. After a time, a solid crust will begin to form on the surface. When the crust is continuous, pierce it at two widely separated points with a glass rod and rapidly pour out the liquid sulphur from inside With a pen-knife, cut through the solid crust all the way round the dish, near the rim and lift it out. Underneath will be seen long needle-shaped crystals of monoclinic sulphur.

X B Preparation of plastic sulphur

Materials required

Test tube	Sulphur	5 g
Beaker (250 ml)	Cold water	

Procedure

Heat powdered sulphur in a test tube until it is boiling rapidly. Then pour the boiling sulphur in a thin continuous stream into a beaker full of cold water. It forms long, elastic, light yellow ribbons of 'plastic sulphur'.

IX C Solubility of sulphur

Materials required

Test tube	Carbon disulphide	5 ml
Evaporating dish	Sulphur	1 g

Procedure

Take about 1 g of sulphur in a test tube Add 5 ml of carbon disulphide to this and shake well. Sulphur dissolves completely.

Transfer the clear solution to an evaporating dish and set it aside for about ten minutes. Evaporation of carbon disulphide takes place and residue of sulphur is left in the dish. Examine the residue with a lens. Crystals of rhombic sulphur can be seen.

IX D Reaction between sulphur and copper turnings

Materials required

Hard glass test tube	Sulphur	1 g
Test tube	Copper turnings	1 g
Test tube holder	Dil. hydrochloric acid	2 ml

Procedure

Take 1 g of sulphur and 1 g of copper turnings in a hard glass test tube. Using the test tube holder, heat the test tube for about 5 minutes. A black substance is formed. Cool the test tube. Transfer the contents of the test tube on to a piece of paper. Take the black product in a test tube. Add about 2 ml of dilute hydrochloric acid and warm(if necessary, heat). A gas with the smell of rotten eggs is evolved, indicating that it is hydrogen sulphide

Elicit from the pupils that the black substance is copper sulphide.

Discussion

Copper reacts with sulphur to form black cuprous sulphide. On adding dilute hydrochloric acid and warming, hydrogen sulphide (a gas with a foul smell) is liberated.

$$2\,Cu + S \longrightarrow Cu_2S$$

$$Cu_2S + 4HCl + O \longrightarrow 2\,CuCl_2 + H_2O + H_2S\uparrow$$

BACKGROUND INFORMATION

6.1 Air

Air is a mixture and not a compound for the following reasons:

1. The nitrogen and oxygen (and other gases) can be separated from each other by physical means. For example, air can be liquified and from liquid air oxygen and nitrogen can be separated by fractional distillation. Water dissolves air and the gas obtained by boiling the water contains more oxygen than air does. In both cases, the oxygen and nitrogen are partly separated by physical means; this could not be done if air was a compound.
2. The properties of air are just the average of the properties of nitrogen and oxygen. For example, the density of air (14.4) is between that of nitrogen (14) and oxygen (16). Oxygen supports combustion very well, nitrogen does not; air supports combustion fairly well
3. The proportions of nitrogen and oxygen (and other gases) in air are not definite; they vary from place to place. Compounds always have a constant composition.
4. Air can be made by mixing nitrogen and oxygen (and small amounts of other gases). No heat or light or any other form of energy is given out or absorbed. In the formation of a compound energy is generally released or absorbed

Composition of dry air by volume

Nitrogen	78%
Oxygen	21%
Carbon dioxide	0 04%
Rare gases	0.96%

Note:

Besides the above, air contains a small amount of water vapour, which is variable. It also contains certain impurities like hydrogen sulphide, sulphur dioxide, dust particles and bacteria.

6.2 Oxygen

Chemical symbol	O	Atomic weight	16
Molecular formula	O_2	Molecular weight	32

Discovery

Oxygen was discovered by Scheele (1771) and by Priestley independently. Lavoisier gave it the name oxygen, meaning "born from acid".

The density of oxygen at 0°C and 760 mm pressure is 1.43 g per litre. It can be liquified to a blue liquid (b.p.-183°C) and frozen to a blue solid.

Oxidation

Many substances burn in oxygen. This process was originally termed oxidation. This term now indicates addition of oxygen or elements like chlorine, iodine etc., or removal of hydrogen or a metallic element.

According to electronic theory, loss of one or more electrons by an atom or ion is oxidation

Reduction

Reduction is a reaction which involves addition of hydrogen or any other metallic element; or removal of oxygen or chlorine, iodine etc.

According to electronic theory gain of one or more electrons by an atom or an ion is reduction.

The processes of oxidation and reduction always occur simultaneously, so that the electrons lost by one atom or ion are taken up by the other atom or ion

Oxidising agents

An oxidising agent is a substance which brings about the oxidation of various substances and in the process it gets reduced. Oxygen, chlorine, bromine and iodine are some of the oxidising agents among elements while nitric acid, potassium dichromate, potassium permanganate and various peroxides are common examples of oxidising agents among compounds.

Reducing agents

A reducing agent is a substance which brings about the reduction of various substances and in the process it gets oxidised. Hydrogen, carbon, sodium potassium, calcium, etc., are some of the reducing agents among elements while hydrogen sulphide, hydrogen iodide, carbon monoxide, sulphur dioxide, ferrous sulphate, stannous chloride, etc., are some common examples of reducing agents among compounds

Combustion is defined as any chemical action in which heat and sometimes light are produced. Breathing and rusting are slow combustion processes. Combustion may or may not involve oxygen. Thus combustion takes place when hydrogen burns in chlorine. No oxygen is involved here. The chlorine acts as the "atmosphere" that supports the burning of hydrogen.

The rate of combustion is determined by several factors .

i. the size of the particles of the substance
ii the amount of oxygen or other "atmosphere" available; and
iii. the nature of the substance.

Catalysis

It has been found that on the addition of some manganese dioxide to potassium chlorate, the latter liberates oxygen faster and at a lower temperature of about 200°C. Manganese dioxide remains unchanged chemically during the experiment. A substance like manganese dioxide that alters the rate of a chemical reaction, itself remaining unchanged is called a catalyst or catalytic agent Other examples are: finely divided platinum in the combination of oxygen with sulphur dioxide to form sulphur trioxide and oxides of nitrogen in the combination of sulphur dioxide, oxygen and steam to form sulphuric acid The last two processes are used in the manufacture of sulphuric acid.

A catalyst cannot affect the total amount of substance decomposed. It merely changes the rate of the reaction; catalysts which increase the reaction rate are more often used than those which retard a reaction.

Role of oxygen in nature

All living beings need oxygen , breathing is not possible without this gas. Oxygen is necessary for burning and rusting.

Free oxygen is present in the air, upto 21% by volume (or 23% by weight) Almost half of the weight of the earth's crust consists of oxygen in a combined state in the form of water, silicates and many metallic and non-metallic oxides.

Fish and other aquatic animals breathe in the oxygen that is dissolved in water. Plants also inhale oxygen. The gas also takes part in the formation of soil It participates in the decomposition of animal and vegetable remains. New substances formed in this process remain in the soil and are taken up by the plants for their nourishment. Thus we see the role of oxygen in various processes of transformation of substances in nature

6.3 Nitrogen

Chemical symbol	N	Atomic weight	14
Molecular formula	N_2	Molecular weight	28

Discovery

In 1772 Scheele put forward the idea that air is a mixture of two gases, one of which supports combustion (fire air) and the other does not (foul air). Daniel Rutherford prepared nitrogen in 1772 by removing the oxygen of the air by combustion and removing the CO_2 so formed by dissolving it in alkali solution. In 1775 Lavoisier recognised nitrogen as an element and called it azote (since it does not support life).

Nitrogen condenses to a liquid at -196°C and freezes to a solid at -210°C. Nitrogen is absorbed by charcoal if cooled to liquid air temperature.

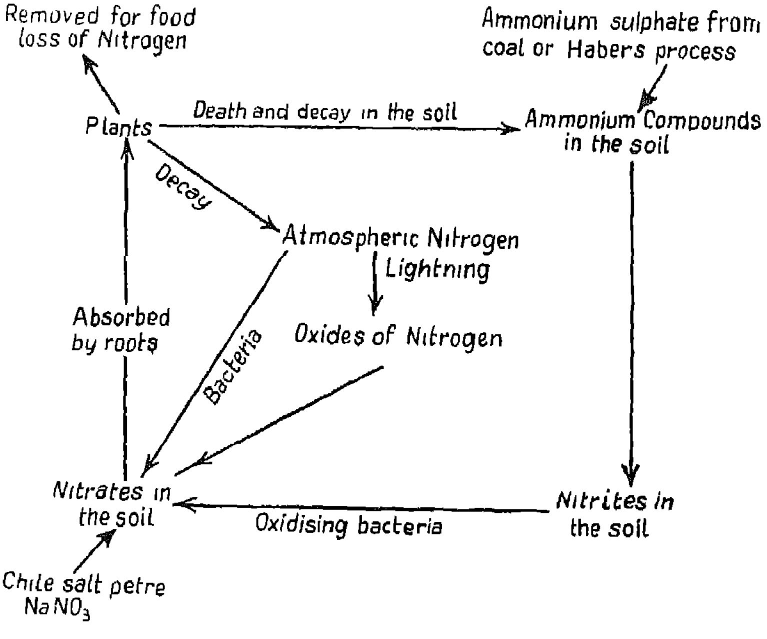
Fixation of atmospheric nitrogen and the nitrogen cycle

Nitrogen forms nearly four-fifths by volume of the atmosphere. This is almost an unlimited reservoir of nitrogen, but its conversion to useful nitrogen compounds is difficult. The soil usually contains combined nitrogen in the form of nitrites, nitrates and ammonium compounds. The tissues of all living organisms, both animal and plant, contain combined nitrogen in compounds known as proteins. Nitrogen in the combined state also occurs to some extent in nature as sodium nitrate which occurs in large quantities in Chile (South America), and is commonly called Chile salt petre. Organic nitrogen compounds are found in all living matter in the form of plant or animal proteins.

The fertility of soil depends in part on the presence in the soil of certain chemical elements K, P, Fe, S and N. Among these elements, nitrogen is very important.

As plants and crops grow from a given patch of soil, some of the nitrogen previously contained in the soil is removed in the form of complex organic compounds, to form parts of the tissue of the plant. It is obvious that unless nitrogen is continually supplied to the soil to balance the loss suffered the fertility of the soil will decrease and the crop yield falls.

The soil receives some nitrogen by natural means. Certain plants, for example, peas and beans, always have colonies of bacteria on their roots which are able to convert the nitrogen of the air into compounds which pass into the soil. Electrical discharges in the atmosphere, such as lightning, cause some slight combination of oxygen and nitrogen and this leads to the passage of nitrogen into the soil as nitric acid dissolved in rain water. Nitrogen fixing bacteria, living free in the soil, are another important agency supplying nitrogen to the soil from the air.



Nitrogen cycle

Fertilizers

Ammonium sulphate, sodium nitrate and farmyard manure are used as nitrogenous fertilizers, to make good the loss of nitrogen.

Explosives

Some of the nitrogenous compounds like nitroglycerine, gun-cotton and trinitrotoluene (T. N. T.) are used as explosives.

6.4 Carbon dioxide

Carbon dioxide is readily liquified by pressure or strong cooling and the liquid mixed with ether forms a freezing mixture, with which a temperature of -100°C may be reached. Carbon dioxide is appreciably soluble in water dissolving in its own volume of water at 15°C under normal pressure.

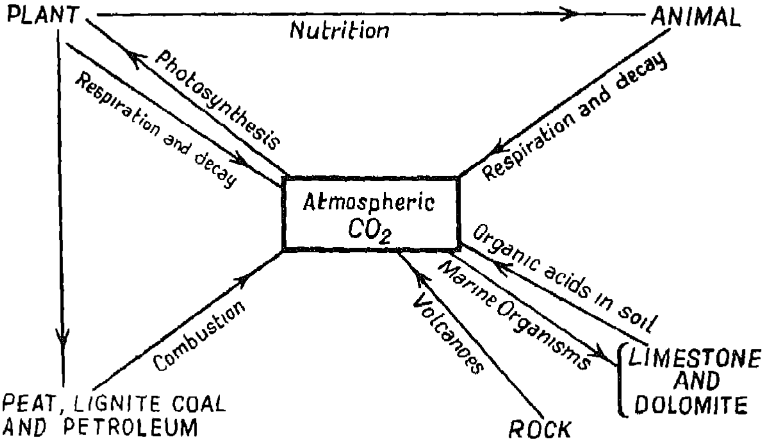
Cycle of carbon and carbon dioxide in nature

Carbon dioxide occurs in the atmosphere to the extent of 0.03 to 0 04 per cent by volume, though the amount varies from place to place This is a fact of utmost importance and significance for, upon the presence of this carbon dioxide depends the continued existence of life. Both plants and animals consist essentially of carbon compounds; and the source of the carbon contained in them is ultimately the carbon dioxide in the atmosphere. Plants in sunlight are able to bring about what is known as photosynthesis, that is they can build up sugars, starches and other organic compounds from the carbon dioxide of the atmosphere and water. The carbon compounds contained in animals are derived from plants eaten by animals.

The carbon thus absorbed by animals is ultimately partly returned to the atmosphere in the form of carbon dioxide exhaled by the lungs, while the decomposition of the bodies of plants and animals also liberates carbon dioxide. There is thus a balance maintained between plant and animal life which results in the maintenance of the proportion of carbon dioxide in the atmosphere at a remarkably steady level.

There are places near the volcanic regions of Naples and Java where large quantities of carbon dioxide issue out from the earth. Carbon dioxide is appreciably soluble in water and a certain amount of carbon dioxide is removed from the air through the formation of shells composed of calcium carbonate by aquatic animals. This is the origin of the enormous beds of limestone and chalk which exist in various parts of the world. Against this may be set the fact that enormous quantities of coal and oil are converted into carbon dioxide during burning.

Carbon passes through a cycle of changes in nature through carbon dioxide of the atmosphere. It resembles the cycle of nitrogen. Both these cycles are essential for the maintenance of life processes on the earth. The cycle of carbon in nature is represented below.



Cycle of carbon and carbon dioxide in nature

Fire extinguishers

Because carbon dioxide suppresses combustion, it is used as a fire extinguisher of which several types are known In one type, sodium carbonate solution is kept apart from an acid by a glass partition. In the case of fire, the glass partition is broken when carbon dioxide is produced under sufficient pressure to force it out along with a liquid charged with the gas, both of which together readily smother the fire.

Production of natural water for drinking purposes

Naturally occurring water contains impurities dissolved from the rocks and soil. Even rain water contains particles of dust and gases dissolved from air. The supply of pure water for drinking purposes is important to the health of a community. Many diseases are water-borne and drinking water must therefore, be completely free from the bacteria causing them. The impurities commonly present in natural water may be classified as follows .

a. Suspended matter of all kinds, generally clay, mud, leaves, etc.

b. Dissolved solids, which usually consist of mineral matter dissolved by water on passing through the ground.

c. Dissolved gases, mostly air.

d. Bacteria, which are found on particles of suspended matter.

The commonest method for the purification of water for town supply (when the supply is derived from a river or canal) is as follows :

The raw water containing suspended impurities (clay, bacteria) is first run into a settling tank. A part of the impurities settles to the bottom by gravity. In some installations the requisite amount of filter alum (arsenicfree alum) is added. This treatment coagulates (brings together) the finely divided mud in the water and thus hastens the settling.

The partly clarified water from the settling tank is run off through an outlet to the sand filters (filter beds). This consist of a top layer of fine sand, a middle layer of coarse sand and a bottom layer of graded gravel. Sand filter acts as the filtering medium and holds up almost all suspended matter including some of the bacteria. The filtered water is completely free from suspended impurities It is not quite safe for drinking purposes as it still contains some harmful bacteria. The water must, therefore, be sterilized before use. Chlorine is commonly used as the sterilizer. The requisite amount of this gas is bubbled into the water. Sometimes chlorine is used either as liquid chlorine or in the form of bleaching powder. A mixture of chlorine and ammonia is found to be more efficient in killing bacteria. Ozone is sometimes employed to sterilize water. Exposure to ultraviolet light (provided by mercury arc lamps) sterilizes water most effectively. The last two methods are preferred to chlorine treatment as no taste or odour is imparted to the water, but these methods are more expensive than chlorination. Sterilized water is then distributed through pipes.

6.6 Hydrogen

Chemical symbol	H	Atomic weight	1.008
Molecular formula	H_2	Molecular weight	2 016

Discovery

As early as the 16th century it was observed that the action of sulphuric acid on certain metals leads to the formation of a gaseous inflammable substance. Because of its ability to burn, it was named 'inflammable air'. In 1776 the 'inflammable air' was shown by Cavendish to be a new substance as it produced water on burning. It was therefore given the name hydrogen (water producer) by Lavoisier.

Properties

Hydrogen liquifies at a very low temperature -252.8°C and at -259.2°C it turns into a snow like solid mass.

Hydrogen is the lightest substance known. It weighs only 0.089 g per litre under normal conditions. It is sparingly soluble in water, 1 ml of water dissolves 0.0215 ml of the gas at 0°C.

The pure gas is not poisonous, but does not support life.

Nascent hydrogen

Gaseous hydrogen at ordinary temperature and pressure is not very reactive. It is considerably more active the moment it is set free in an exorthermic reaction, and when it is in close proximity with the molecules on which it is going to react. It is then said to be in the **nascent state** If gaseous hydrogen is passed through a solution of ferric chloride no chemical reaction occurs. If, however, zinc and hydrochloric acid are added to a reddish-brown ferric chloride solution, hydrogen is produced in the solution and is then capable of reducing the ferric chloride to pale green ferrous chloride.

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + 2H \text{ (nascent)}$$

$$\underset{\text{(Reddish-brown)}}{FeCl_3} + H \longrightarrow \underset{\text{(Pale green)}}{FeCl_2} + HCl$$

Similarly nascent hydrogen reacts with acidified $KMnO_4$ solution and reduces it to potassium sulphate and manganese sulphate, the solution of these salts being almost colourless; the nascent hydrogen is itself oxidised to water.

Hydrogenation of oils

Oils are changed to solid fats when hydrogen gas is passed into them under pressure and in the presence of finely divided nickel used as catalyst.

$$\text{Vegetable oil} + H_2 \xrightarrow[\text{Catalyst}]{Ni} \text{Fat}$$

This process is used in the preparation of an **artificial ghee** or "Vanaspati".

6.7 Common salt or sodium chloride

Formula NaCl

Formula weight 58.6

Sea water is the richest source for common salt. In India, common salt is produced by pumping sea water into large, open fields and allowing the water to evaporate by the heat of the sun.

The common salt thus prepared contains traces of magnesium chloride and calcium chloride. Since these salts are hygroscopic, i. e., absorb moisture from air, the commercial common salt becomes hygroscopic.

The commercial salt is purified by recrystallisation (table salt) or by passing hydrogen chloride gas through a saturated solution of the common salt.

The presence of traces of magnesium salts is very essential in our food So, very pure sodium chloride is not recommended as table salt.

6.8 Chlorine

The elements fluorine, chlorine, bromine and iodine form a closely related family of elements known as the halogens. The term halogen means "salt former" since all of these elements occur in sea-water and form compounds resembling common salt.

Chlorine was discovered by the Swedish chemist Scheele in 1774, who obtained it by heating manganese dioxide and hydrochloric acid. It was in 1809 that the nature of chlorine as a chemical element was proved by Davy.

Physical properties

The poisonous nature of chlorine is well known since its use in World War 1 It is about 2 5 times as dense as air and is moderately soluble in water. The resulting yellowish solution is known as chlorine water. Chlorine is readily liquified at room temperature by a pressure of about 6 atmospheres. It is then stored in steel cylinders.

Chemical properties

Chlorine as an oxidising agent

When a gas jar of hydrogen sulphide is inverted over a gas jar of chlorine, yellow precipitate of sulphur and hydrogen chloride are formed.

$$H_2S + Cl_2 \longrightarrow 2\,HCl + S\downarrow$$

This is due to the combination of chlorine with the hydrogen of hydrogen sulphide forming hydrogen chloride and sulphur It is an example of oxidation and reduction.

Bleaching action of chlorine

Chlorine in the presence of moisture bleaches a large number of coloured organic compounds. The bleaching action is due to the hypochlorous acid (HOCl) formed by the action of chlorine on water. Hypochlorous acid is unstable and loses oxygen. The nascent oxygen so liberated bleaches the organic substance.

$$Cl_2 + H_2O \longrightarrow HOCl + HCl$$

$$\underset{\text{Coloured}}{\text{Dye}} + HOCl \longrightarrow HCl + \underset{\text{(Colourless)}}{(\text{dye} + O)}$$

From the equation above, it is seen that water is necessary for the formation of HOCl, the compound which liberates the oxygen and which performs the bleaching. Hence, if no water is present no bleaching can occur. Thus chlorine bleaches by oxidation in the presence of moisture.

Notice that hydrochloric acid is produced whenever chlorine bleaches, hence the articles must be thoroughly washed after bleaching or it will be attacked by the free acid.

Bleaching powder

If chlorine is passed for a considerable time over solid slaked lime, the product is bleaching powder It contains about 36% of available chlorine, i. e., chlorine which can be removed by dilute acids and even by carbon dioxide of the atmosphere. Hence bleaching powder usually smells of chlorine and deteriorates in contact with air. It has an extensive use in dye works, laundry and even in the home.

6.9 Carbon

Allotropy

If an element exists in more than one 'form' then this property is known as 'allotropy' and the different forms are called 'Allotropic modifications'. The different forms of the same element possess different physical properties but have similar chemical properties. Allotropy is shown by metals as well as by non-metals, e g., tin, iron, oxygen, carbon, sulphur and phosphorus. In the case of carbon, allotropy is due to the existence of two or more different crystalline forms, e. g., diamond and graphite are two varieties of carbon which have different crystalline forms

Allotropic forms of carbon

The element carbon exists in two different forms — diamond and graphite. The other forms, e. g., charcoal, coke, soot are very tiny crystals of graphite.

The forms of carbon differ in appearance and in many other physical and chemical properties. When 1 g of diamond or graphite is burnt in oxygen, it forms 3.67 g of carbon dioxide and nothing else Charcoal is not quite pure carbon. Diamond and graphite are different because the carbon atoms are arranged differently in their crystals.

In diamond each carbon atom is surrounded by four other carbon atoms, to form a tetrahedron. A tetrahedron is a four sided body In other words, each carbon atom acts as the centre of a tetrahedron in which it is surrounded by four other carbon atoms, each of which in turn acts as the centre of another tetrahedron in which it is surrounded by four other carbon atoms. This arrangement produces a characteristic three-diamensional structure for diamond. As a consequence of this three diamensional structure, a given piece of diamond can be regarded as a giant molecule in which each carbom atom (except those on the surface of the crystal) is joined by single covalent bonds to four other carbon atoms. Since the bonds are of equal strength and very strong, diamond is a very hard substance and has a very high melting point

In graphite also the carbon atoms are arranged according to a regular pattern. In contrast to the structure of diamond, however, the carbon atoms in graphite are bonded strongly to one another only in one plane, and more weakly in a three-dimensional pattern. Each carbon atom in graphite is linked to three other carbon atoms to form a two-dimensional net work made up of hexagons of carbon atoms. That is, in graphite, the carbon atoms are arranged in layers of hexagons The bond between the layers is weak. The layers can slide over one another easily. That is why graphite is flaky or greasy and is used as a lubricant.

Diamonds are mined in South Africa and also in India and Brazil. Graphite is mined in Ceylon and is also made artificially by heating coke in an electric furnace to a very high temperature

The following table compares the properties of the three forms of carbon :

Property	Diamond	Graphite	Charcoal
Appearance	Colourless, transparent, sparkling	Black, opaque, shiny	Black, opaque, dull
Sp. gravity	3.5	2.3	1.5 when free from air
Hardness	Hardest substance known	Soft, greasy	Soft
Burning in air	Least readily at 900°C	At about 700°C	Most readily at about 500°C
Electrical conductivity	Non conductor	Good conductor	

Diamonds are valuable gems. They are used for cutting glass and rock The lead of pencils is graphite mixed with clay. Graphite is used as a lubricant for electric motors and fast moving parts of machinery. It is also used for making electrodes because it is a conductor of electricity and is not affected by most substances.

Other forms of impure carbon

1 *Coke*

It is the solid left when coal is heated out of contact with air (destructive distillation of coal). Coke is used as a fuel and also as a reducing agent in the extraction of certain metals.

2. *Wood charcoal*

It is the solid left when wood is heated out of contact with air. Wood charcoal is a black porous solid. It is used as a domestic fuel. A piece of wood charcoal can absorb about 100 times its volume of some gases. It is therefore used in gas masks.

3. *Animal charcoal*

It is the solid left when bones are heated out of contact with air. It contains about 10% of carbon and the rest is mainly calcium phosphate. It absorbs the colouring matter from solutions and is therefore used the refining of sugar.

4. *Lamp black*

It is formed when paraffin oil or turpentine burns in a limited amount of air. It is used to make Indian ink, black shoe polish, carbon paper and printer's ink.

5. *Soot*

It is formed by the incomplete combustion of coal or other fuels.

Coal

Coal is a complex mixture of compounds of carbon, hydrogen and oxygen together with small amounts of compounds of nitrogen, sulphur and phosphorus. It contains 30% to 90% of carbon. It 'was probably formed thousands of years ago from forests. The trees and plants of the forests died and were slowly burried by soil, sand and mud. The wood decayed under great pressure and gradually changed to coal.

Coal is used as a fuel. It is also used to make coal-gas and as a fuel in generating steam and electricity.

Unique nature of carbon

The element carbon possesses certain unique properties. The most important of these properties is the unique ability of carbon atoms to combine with one another to form "chains" which may be straight, branched or cyclic. Consequently, an infinite number of carbon atoms can unite with one another, the first with the second, the second with the third, and so on to form molecules that consist of carbon atom chains of different lengths. No other element is capable of forming such a range of stable atom-to-atom links.

Another important property of the element carbon is that it is capable of combining with many other elements such as oxygen, hydrogen, sulphur, nitrogen and halogens and other non-metals. It can even combine with the atoms of metals.

Although carbon is not like the alkali metals, nor like the halogens, it is known to form linkages with elements of both these extreme types.

Because of these unique properties of carbon, compounds of carbon are more numerous than the compounds of all the other elements put together. There is a separate branch of chemistry dealing with carbon compounds called organic chemistry.

6.10 Sulphur

Allotropic modifications of sulphur

The allotropes of sulphur are divided into (1) crystalline and (2) amorphous or non crystalline.

The important crystalline forms are:

a) Rhombic or α–sulphur,
b) Monoclinic or prismatic or β-sulphur.

The important amorphous forms are:

a) Plastic sulphur; and
b) Milk of sulphur.

Crystalline forms

a) *Rhombic sulphur*

This is the most stable variety of sulphur at ordinary temperature and all other forms pass into it on standing. To prepare rhombic sulphur a sample of sulphur is ground up and shaken with carbon disulphide and filtered. The filtrate is allowed to evaporate slowly in an evaporating dish. Transparent lemon-yellow rhombic crystals of sulphur separate

Rhombic sulphur has a sp. gravity of 2.06, m. p 112.8°C and b. p. 445°C. It is insoluble in water but soluble in carbon disulphide, hot benzene, carbon tetrachloride and turpentine. It is stable below 96°C.

b) *Prismatic sulphur*

It is prepared by melting sulphur in a crucible and allowing it to cool slowly till a solid crust forms on the surface.

The crust is then pierced at two points and the liquid sulphur poured away. On removing the surface crust, transparent long needle shaped yellow crystals of prismatic sulphur will be found inside the crucible.

It has a sp. gravity of 1.96 and m. p. 119.25°C. It is insoluble in water but readily soluble in CS_2. It is stable between 96° and 119°C. When cooled below 96°C it gradually changes to rhombic sulphur.

Amorphous form

Plastic sulphur

This is prepared by heating sulphur to its boiling point and then pouring the black mobile liquid into cold water, when it solidifies to viscous elastic transparent yellow strings of plastic sulphur.

It has a sp. gravity 1.92. It is insoluble in CS_2. It is so viscous that it may be drawn out into long threads. On standing, however, it gradually changes to a hard brittle mass, which is in part rhombic sulphur.

Properties of sulphur

The effect of heat on sulphur is interesting. On heating it melts to an amber coloured liquid As the temperature is raised, it becomes much darker in colour and viscous So viscous does it become that the test tube may be inverted without the sulphur flowing. On further heating, the liquid retains its colour slowly, loses its viscosity and finally boils at 445°C, giving of vapours of sulphur. On cooling the boiling liquid slowly, these changes take place in the reverse order till the sulphur is obtained as a pale yellow solid.

Quiz

1. Tick off (✓) the correct answer in the following.
   - i. A substance which gives oxygen on heating is
     - a. potassium permanganate
     - b. ammonium nitrite
     - c. ammonium chloride.
   - ii. A substance which gives nitrogen on heating is
     - a. ammonium nitrite
     - b. potassium chloride
     - c. lead nitrate.
   - iii. A reaction in which hydrogen is evolved is
     - a Calcium carbonate + hydrochloric acid
     - b. Cupric oxide + carbon
     - c. Zinc + dilute sulphuric acid.
2. Write in one word what the following describe .
   - a Substance formed when ammonia and hydrogen chloride react.
   - b. Precipitate obtained on passing carbon dioxide into lime water.
   - c. Substance formed when sulphur burns in oxygen.

3. State whether the following statements are True or False.
   - a. Diamond and graphite are forms of carbon.
   - b. Carbon dioxide is a bleaching agent.
   - c. Manganese dioxide acts as catalyst in the liberation of chlorine from hydrochloric acid.
4. Fill up the blanks with suitable words.
   - a. Chlorine is a .. .. coloured gas
   - b. Carbon dioxide is .. .. than air.
   - c. .... burns in nitrogen.
5. Explain why
   - a. when ice pieces are kept inside a glass, water droplets are formed on the outside surface
   - b. carbon dioxide is used as a fire extinguisher
   - c. mountaineers carry oxygen
   - d. hydrogen is collected over water.
6. A certain gas extinguished a candle flame immediately; it had no action on lime water Which one of the following could it be?
   - a Air
   - b. Nitrogen
   - c. Carbon dioxide
   - d. Oxygen.

7

Common Metals and Alloys

INSTRUCTION MATERIAL

Introduction

About three-fourths of the known elements are metals. They are widely used in daily life. Metals occupy a prominent position in any civilized economy. It is difficult to imagine the present day world without metals. The specific uses to which they are put depend on their characteristic properties. Most of the metals are rare and only about twenty of them are commonly used in daily life.

A few metals such as gold are found free in nature. But most of them are not and so have to be extracted from their ores which are minerals containing metal compounds. These minerals usually occur along with other materials. It is to the great credit of the metallurgical chemist that so many metals have been extracted and put to various uses for the benefit of mankind. Of late, intensive research is being carried out in finding newer uses for the rare metals such as beryllium, vanadium and molybdenum.

Outline

1. How do we recognise metals? (7.1).
2. Reactivity of metals (7.2).
3. How do metals occur in nature? (7.3).
4. Obtaining a metal from its ore (7.4)
5. Alloys (7.5).

Concepts

1. Metals are recognised by their lustre, malleability, ductility and conductivity.

2. Metals vary in their reactivity towards air and water. Many metals of everyday use corrode in moist air and this can be prevented in different ways.
3. An alloy may be obtained by heating molten metals together. Alloying modifies the properties of a metal.

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	27	Discussion	I and II A	Importance of metals and 7.1
2	—	—	II B	7.2 a
3	—	—	III and IV	7.2 b, c and 7.3
4	28	Discussion	V	7.4
5	—	—	VI	7.5
6	—	—	—	Review and quiz

Development

The importance of metals in daily life may be impressed upon the pupils through demonstration I.

Demonstration I Metals in everyday life

Time required · 15 minutes

Materials required

Cooking vessels
Door handles of iron, brass or aluminium
Nails and bolts
School bell
Files, spanners and cycle parts
Copper and iron wires
Sheets of copper, aluminium, iron, tin, etc.
Thermometer
Electric bulb
Chocolate wrappers and seals of milk bottles

Procedure

Show the above articles one by one to the pupils. Ask simple questions to make them realise that

i. cooking vessels are made of copper, brass, aluminium, or stainless steel

ii. common tools and appliances are made of iron, steel, brass, aluminium, etc.

iii. mercury is used in thermometers

iv. filament of an electric bulb is made of tungsten

v. aluminium foils are used as wrappers and seals.

In the course of the demonstration, make the pupils realise that living conditions have considerably improved by the use of metals. Select such articles as are familiar to the pupils. The list cited above is only indicative.

It is advisable not to classify elements into metals and non-metals, or give formal definition of metals. The various uses to which the metals are put in daily life need not be listed.

7.1 How do we recognise metals ?

Experiment 27 is designed to make the pupils understand some of the properties of metals such as lustre, metallic sound, malleability, thermal conductivity and electrical conductivity.

Experiment 27 Properties of metals

Time required : 30 minutes

Questions and answers

1. What is the effect of rubbing metals with sand paper ?
   The metal surface becomes bright and shining.

2. Why are some metal surfaces dull before rubbing ?
   The dullness of the metal surface is due to the coating formed by the action of air and moisture on the metal.

3. What does the glow of the bulb indicate ?
   The glow indicates that current is flowing through the metal strip.

Metals are lustrous because they are good reflectors of light. The shining surfaces of gold, silver and stainless steel may be cited as familiar examples in addition to the metals studied. Mirrors are made by coating glass with a thin layer of metal.

Metals are malleable, that is, they can be twisted, bent or spread into sheets. Gold and silver may be beaten into very thin foils Lead, tin, zinc and copper are also malleable. Some metals, however, are brittle.

Metals are ductile, and therefore can be drawn into thin wires with the help of dies. Platinum and gold are drawn into very fine wires.

Metals are good conductors of heat and electricity Of all the metals, silver has the highest thermal and electrical conductivities. However, because of low cost, aluminium and copper are used in making water boilers and electrical cables.

All metals, with the exception of mercury, are solids at ordinary temperature. Most metals have rather high melting points. Metals vary in hardness.

7 2 Reactivity of metals

Metals vary in their reactivity towards air. Demonstration II has been designed to enable the pupils to see the difference in the reactivity of sodium, magnesium, aluminium, iron and copper towards air. The demonstration is divided into two parts. II A deals with action of air at ordinary temperatures and II B, at higher temperatures.

Demonstration II Action of air and oxygen on metals

Time required · 50 minutes

II A Action of air on metals at ordinary temperatures

Materials required

Glass plate (10 cm × 10 cm)	Sodium metal*	
Sand paper	Magnesium ribbon	2 cm
Watch glass	Aluminium sheet	1 sq cm
	Iron strip	1 sq cm
	Copper foil	1 sq cm
	Silver foil	1 sq cm

Procedure

Cut a small piece of sodium and place it on a watch glass. Ask the pupils to observe it carefully for some time. Let them observe that the bluish white shining surface of the freshly cut piece tarnishes soon.

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*For precautions while handling sodium, refer unit 1, page 15

Clean the samples of other metals by rubbing with sand paper and place them on a clean dry glass plate. Set aside the glass plate. Ask the pupils to observe the samples everyday for one week

II B *Action of air and oxygen on metals at high temperatures*

Materials required

Spirit lamp	Sodium metal
Tongs	Aluminium powder
Box of matches	Iron wire
Sand paper	Silver wire
	Jars of oxygen 3

Procedure

Cut a small piece of sodium and direct the attention of the pupils to its appearance. Then heat it on a deflagrating spoon. Let them observe that sodium melts and burns with a yellow flame forming a light yellow product (sodium peroxide)

Recall from unit 5, the experiment on burning of magnesium in air, and from unit 6, the demonstration on burning of magnesium in oxygen.

Heat aluminium powder in a deflagrating spoon strongly and point out that it does not burn in air. Then introduce the hot powder into a jar of oxygen. Show that aluminium burns now forming a white powder (aluminium oxide).

Recall from unit 6, the demonstration on burning of copper wire in oxygen.

Take iron wire and rub it with sand paper until it is shining bright. Then heat it till it is red hot and introduce it into a jar of oxygen. Show that it burns and forms bluish black substance (iron oxide).

Take the clean silver wire and heat it until red hot. Introduce it into a jar of oxygen Show that it remains unaffected.

The above demonstrations illustrate that

i. sodium readily reacts with air at room temperature whereas magnesium, aluminium, iron and copper react slowly.

ii. sodium and magnesium burn in air when heated. The other metals, namely, aluminium, copper and iron react with oxygen only when strongly heated or when red hot.

iii. silver is unaffected by air at room temperature or even by oxygen at high temperatures.

Thus it is seen that metals differ in their reactivity towards oxygen.

Demonstration III gives further information about the reactivity of metals.

Demonstration III Action of water on metals

Time required . 40 minutes

Materials required

Beaker (400 ml)		Sodium metal	
Test tubes	2	Magnesium ribbon	3 cm
Pair of tongs			
Corks for test tubes			
Box of matches			
Splinter			
Funnel (5 cm dia) with a short stem			
Copper gauze			

(if copper gauze is not available, wrap the piece of sodium loosely in a piece of cigarette wrapper).

Procedure

a. Take water in a beaker and invert the funnel in it such that the stem is under water. Fill a test tube completely with water and invert it carefully over the stem of the funnel. Cut a small piece of sodium, wrap it in a piece of copper gauze and introduce it carefully under the funnel. Let the pupils observe the evolution of a gas. Test the gas with a burning splinter to show that it is hydrogen.

b. Rub a magnesium ribbon with sand paper until it is bright. Place it in a beaker containing water. Show that no reaction occurs. Invert a funnel over it such that its stem is under water. Heat the water to boiling. Fill a test tube with water and cork it. Invert it in the boiling water and remove the cork with the help of tongs. Place the test tube over the stem of the funnel. Show that the gas evolved is hydrogen. Thus point out that magnesium which does not seem to react with cold water, reacts with boiling water.

Discussion

The above demonstration illustrates that sodium is more reactive towards water than magnesium. Copper is not affected by cold or boiling water. Silver and gold are not affected even by steam. Thus it is seen that metals differ in their reactivity towards water.

7.3 How do metals occur in nature?

The occurrence of metals in nature can be correlated with their reactivity. Metals which are inactive or less active possibly occur free in nature. Noble metals like gold may be cited as examples. How about the more active metals?

As seen earlier, metals like sodium are too reactive to exist free in nature. In fact, sodium is always kept under kerosene to avoid contact with air. Thus the active metals always exist in the combined state. The occurrence of metals in nature is to be illustrated with the help of the follwing demonstration.

Demonstration IV Ores of metals

Time required : 15 minutes

Materials required

- Cerussite
- Copper pyrites
- Bauxite
- Haematite
- Rock salt
- Copper
- Aluminium
- Iron
- Lead
- Sodium

Procedure

Display the above cited ores* without emphasising their names. Tell the pupils about the important ores occurring in various parts of India and metals obtained from these ores. Exhibit a map of India showing occurrence of important ores. While displaying each ore, show a sample of the pure metal obtained from it and draw the contrast.

Ore	**Metal**
Copper pyrites	Copper
Bauxite	Aluminium
Haematite	Iron
Cerussite	Lead
Rock salt	Sodium

The above list of ores is only indicative. More may be added to the above list to include any metals with which the pupils are familiar.

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* The ores listed are only suggestive. You may replace some of these with the locally available ores.

7.4 Obtaining a metal from its ore

Demonstration IV might set the pupils thinking as to how the metals can be obtained from their ores. This is the proper time for assigning the pupils experiment 28.

Experiment 28 Metal from its oxide

Time required : 40 minutes

Questions and answers

1. How is lead obtained from its oxide ? Write the equation.

   Lead is obtained by the reduction of lead oxide.

$$PbO + C \longrightarrow Pb + CO\uparrow$$

2. What is the role of charcoal in this reaction?

   Charcoal powder removes oxygen from lead oxide and thus reduces it to metal, lead.

By performing the above experiment, the pupils learn how an oxide can be reduced to a metal using charcoal powder. Recall that cupric oxide could also be reduced by charcoal powder (page 40). It should be noted that not all the metals occur as oxides in nature. For example, lead occurs as sulphide (galena) and as carbonate (cerussite).

The following demonstration shows how lead could be obtained from cerussite. Cerussite decomposes on heating to give lead oxide which is reduced to lead by heating with charcoal.

Demonstration V Obtaining a metal from its ore

Time required : 40 minutes

Materials required

Hard glass test tube	Cerussite[*]	0.5 g
Charcoal block with a cavity		
Spirit lamp		
Kerosene lamp		
Blow pipe		

Procedure

Take about 0.5 g of cerussite (lead carbonate) in a hard glass test tube and heat it on a spirit lamp keeping the test tube slightly inclined. Let the pupils note the colour change from white to yellow. Show them that carbon dioxide is evolved leaving lead oxide in the test tube.

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[*] If cerussite is not available, use lead carbonate

Mix the product left (lead oxide) in the test tube with charcoal powder and place it in a cavity on the charcoal block. Using a blow pipe, direct the reducing flame of a kerosene lamp on to the mixture in the cavity. Heat for about ten minutes. Let the pupils observe the formation of small shining globules of lead.

7.5 Alloys

For various reasons, pure metals are not used as such in industry and daily life. They corrode, i.e., are not able to stand wear and tear The properties of metals can be modified by mixing them with other metals, known as alloying Formation of an alloy is shown in the following demonstration.

Demonstration VI Preparation of an alloy

Time required : 60 minutes

Materials required

Crucibles	2	Tin	6 g
Tripods	2	Lead	4 g
Clay pipe triangles	2		
Pair of tongs			
Burners	2		
Pyrex glass rod			

Procedure

Take 6 g of tin in a crucible and heat it to melting. Similarly take 4 g of lead in another crucible and heat it to melting. Transfer molten lead to the crucible containing molten tin carefully. Stir the mixture with a pyrex glass rod and continue heating for about ten minutes. Allow the mass to cool and show the resulting alloy of tin and lead. This alloy is called solder.

Also display articles of common alloys such as brass and stainless steel.

BACKGROUND INFORMATION

Historical

Metals like gold, silver, mercury, copper, tin, iron and lead were known to man from ancient times. Sanskrit literature of vedic age contains many references to metals. Gold and silver ornaments are believed to have been the first metallic products of ancient artisans. Later iron was used in making utensils, weapons and ornaments. Probably copper and bronze were used much earlier.

7.1 How do we recognise metals

Metals are in general characterised by hardness, lustre, malleability, ductility, high specific gravity and ability to conduct heat and electricity. Some of these properties like electrical and thermal conductivity, malleability and ductility can be explained on the basis of electronic structure of metals

A metal atom has fewer electrons in the outermost shell as compared to a non-metal. A metal crystal is a collection of a large number of regularly arranged metal atoms A few electrons of these atoms are not confined to any particular atom. They are delocalised. These electrons can move through the whole crystal without being attached to any particular atom. These delocalised electrons are called conduction electrons. It should be remembered that all the electrons in the outermost shell of the atom may not be conduction electrons. Thus a metal crystal can be considered as a collection of positive ions surrounded by a sea of electrons.

When an electrical potential is applied across a metal plate or wire the electrons tend to shift toward the positive terminal. The movement of these mobile electrons under the applied electrical potential constitutes electric current through the metal. No minimum potential difference is necessary for establishing the flow of current in a solid or liquid metal.

When we apply heat to a particular part of the metal, the nuclei vibrate with a greater frequency (their kinetic energy increases). The vibrating nuclei transfer this kinetic energy to the mobile electrons. The mobile electrons carry this energy to nuclei in other parts of the crystal. This accounts for the thermal conductivity of metals.

The positive ions are able to slip past one another and still remain bound in a crystal because of the delocalised electrons. Continued slippage of ions past each other will allow the metal to be drawn into wires or flattened into sheets without breaking. This explains malleability and ductility.

Some useful data

i *Decreasing order of malleability*

Au, Ag, Cu, Sn, Pb, Zn, Fe.

——————————→

ii. Specific gravities and melting points of some common metals

Metal	**Sp. gr.**	**M. P. °C.**
Lithium	0.53	186.0
Sodium	0.97	97.5
Calcium	1.54	851.0
Magnesium	1.74	651.0
Aluminium	2.70	659.7
Zinc	7.14	419.4
Tin	7.28	231.9
Iron	7.86	1535.0
Copper	8.90	1083.0
Silver	10.50	960.5
Lead	11.34	327.5
Mercury	13.54	–38.9
Gold	19.30	1063.0

Metals which have sp. gr. less than 5 are called light metals.

iii. Relative electrical conductivities (taking copper as 100)

Silver	106
Copper	100
Gold	71
Aluminium	65
Tungsten	31
Zinc	29
Iron	17
Tin	15
Lead	8.3
Mercury	1.8

7.3 How do metals occur in nature?

Some relatively inactive metals like gold, silver and platinum occur free in nature. Active metals occur in nature only in the form of compounds. Naturally occurring metal compounds are impure and contain much rocky material. Ores are naturally occurring metal compounds which have sufficient concentration of the metal so that they can be profitably worked for obtaining the metal. Economic extraction of a metal from its ores varies with individual metals Thus an ore containing 10% iron is worthless while one with

10% copper is worth working. Ore with as small as 0.002% gold is considered a good ore of gold.

Occurrence of some metals

Form	Metals
Oxides	Fe, Sn, Al, U, Ti, etc.
Sulphides	Cu, Fe, Pb, Ni, Zn, Hg, etc.
Carbonates	Ca, Mg, Fe, Cu, Pb, Zn, etc.
Sulphates	Ca, Pb, etc.
Chlorides	Na, K, Ag, Mg, etc.

7.4 Obtaining a metal from its ore

The process of extracting a metal from its ore is called metallurgy. Since the natural ores are often complex and impure, the process of obtaining metals from them also becomes quite complicated Again, metallurgy of each metal presents an individual problem depending upon the properties of the metal. The chief operations, especially for heavy metals, include the following steps.

i. Concentration

The earthy and rocky material (collectively known as gangue) is removed in this step. Froth floatation and magnetic separation are some of the common processes used for the purpose.

ii. Roasting

This is heating the concentrated ore in air which serves to remove volatile impurities as well as to convert carbonates and sulphides into oxides.

$$2\,CuS + 3O_2 \longrightarrow 2\,CuO + 2\,SO_2\uparrow$$

$$PbCO_3 \longrightarrow PbO + CO_2\uparrow$$

$$ZnCO_3 \longrightarrow ZnO + CO_2\uparrow$$

iii. Reduction

Roasted ore is heated with an appropriate reducing agent (carbon, either as coal or coke) to get the metal.

$$PbO + C \longrightarrow Pb + CO\uparrow$$

$$ZnO + C \longrightarrow Zn + CO\uparrow$$

Some oxides are not easily reduced by carbon. In such cases aluminium, being a powerful reducing agent, is used for obtaining metals such as chromium and manganese.

$$Cr_2O_3 + 2\,Al \longrightarrow Al_2O_3 + 2\,Cr$$

iv. Refining

Crude metal obtained on reduction is then purified. Refining process varies with the nature of the metal. Volatile metals like zinc and mercury are purified by distillation. A number of metals such as copper are refined electrolytically.

7.5 Alloys

Alloys are obtained by mixing the molten metals and allowing them to solidify. There are three classes of alloys in which

i the component metals crystallise separately (alloys of copper and lead belong to this class).

ii. the component metals completely intermingle and form solid solutions (copper coins and alloys of gold and silver belong to this class).

iii. the components may react to form definite **compounds** (compounds such as Zn_3Ag_2, CuMg, Cu_5Sn, Na_4Pb, Cu_5Zn_8 etc., are known.)

If one of the component metals is mercury, the alloy is known as an **amalgam.**

COMMON ALLOYS

Name		Composition	Uses
Brass	i	Cu 90%, Zn 10%	Cheap jewellery, gold paint.
	ii.	Cu 67%, Zn 33%	Tubes, sheets, etc.
Monel metal		Ni 72%, Cu 26.5%, Fe 1.5%	Non-corrodible propeller blades, wire.
Constantan		Cu 60%, Ni 40%	Thermocouples
Duralumin		Al 95.5%, Cu 3%, Mn 1% Mg 0 5%	Aeroplane and automobile parts.
Magnalium		Al 90%, Mg 10%	Balance beams and light instruments.
Gun metal		Cu 90%, Sn 10%	Gears and castings.
Bell metal		Cu 78%, Sn 22%	Bells, gongs, etc.
Type metal		Pb 82%, Sb 15%, Sn 3%	Casting type.
Bronze		Cu 80-90%, Sn 5-18%, Zn 2-10%	Machinery

Name	Composition	Uses
German silver	Cu 52%, Zn 26%, Ni 22%	Corrosion-resistant vessels etc.
Nichrome	Ni 60%, Cr 12%,Fe 26%, Mn 2%	Resistance wire
Chromium-nickel steel	Fe 74%, Cr 18%, Ni 8%	Cutlery, ornaments
Chrome-vanadium steel	Fe 97-89%, Cr 2-10% V 15%	Automobile parts
Invar	Fe 64%, Ni 36%	Clocks and instruments
Tungsten steel	Fe 80-70%, W. 10-20%, Cr 2-10%	Lathe tools
Stainless steel	Cr 12-14%, Ni .7%, C 0.15 to 1%, Fe 88-85%	Surgical instruments utensils, ete.

Quiz

1. Fill up the blanks.
   a. Corrosion of metals can be prevented by..... .. . .
   b. Some metals when exposed to air acquire . . . . coating.
   c. Metals are . . .. conductors of heat and ... . conductors of electricity.
2 Tick off the correct answer
   Bronze is an alloy of
   a. copper, tin and zinc
   b. copper and zinc
   c. zinc and tin.
3. If you were given a sample of a solid element, describe some simple tests you would perform to show whether it is a metal.
4. Why is pure gold rarely used for jewellery.
5. It is said that iron rusts more quickly in the presence of water containing salt than it does in the presence of pure water. Describe how you would attempt to show whether this statement is true or false.

8

Acids, Bases & Salts

INSTRUCTION MATERIAL

Introduction

Acids and bases play an important role in controlling chemical reactions, whether carried out in laboratory, industry, agriculture or in biological systems.

The treatment of acids and bases is limited mainly to experimental approach. The concept of acids and bases has been developed through their behaviour towards some coloured substances (indicators), metals and carbonates.

Salts, as products of neutralization of acids and bases, and their importance are also discussed.

Outline

1. Acids and bases (8.1).
2. Formation of acids and bases (8.2).
3. Properties of acids and bases (8 3).
4. Neutralization (8.4).
5. Salts (8.5).

Concepts

1. Acids give H^+ ions in aqueous solutions.
2. Bases give OH^- ions in aqueous solutions.
3. Salts are products of mutual neutralization of acids and bases.

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	29, 30	Discussion	—	8.1
2	—	—	I	8 2
3	31	Discussion	—	8.3 and 8.4
4	32, 33	Discussion	—	8.5
5	—	—	—	Review and quiz

Development

8.1 Acids and bases

The idea of acids and bases may be introduced by allowing the pupils to study the behaviour of the given set of solutions towards litmus and other indicators, metals and carbonates as indicated in experiment 29.

The set of substances to be supplied may consist of two acids, two bases and one neutral substance chosen from the following.

Acids	Bases	Neutral substances
Dilute hydrochloric acid	Dilute sodium hydroxide solution	Distilled water
Dilute sulphuric acid	Dilute ammonium hydroxide solution	Sodium chloride solution
Dilute acetic acid	Lime water	
Lemon juice	Sodium carbonate solution	
Soda water		

Do not reveal the names of the substances supplied. Label them as A, B, C, D and E.

Experiment 29 Investigation of properties of acids and bases

Time required : 30 minutes

Postlab discussion

At the end of the experiment, let the pupils group the substances on the basis of similarity in their behaviour. To enable the pupils to identify which are acids and which are bases, take an acid and show its action on litmus, a metal and a carbonate. Show similarly the action of a base on the above materials. Ask the pupils now to label which of the substances given to them are acids, which are bases and which are neutral.

After this, let the pupils perform experiment 30 to test some household materials for their acidic or basic nature. Ask the students to bring the materials from their houses. They may also be asked to collect samples of soils from different parts of their locality. Investigation of soil acidity is important because the fertility of the soil is adversely affected by acidity.

Experiment 30 Testing some household materials and soils for their acidic or basic nature

Time required : 40 minutes

8.2 Formation of acids and bases

Acids, in general, are formed when some oxides of non-metals react with water and bases are formed when metallic oxides react with water. This may be shown by demonstration I.

Demonstration I Formation of acids and bases

Time required : 15 minutes

Materials required

Gas jars	4	Blue litmus papers	4
Deflagrating spoon		Red litmus papers	4
Pair of tongs		Sulphur	0.5 g
Box of matches		Red phosphorus	0.5 g
Spirit lamp		Magnesium ribbon	4 cm
Glass covers (for the jars)			

Procedure

Collect oxygen gas in four gas jars. Take about 0.5 g of sulphur powder in a deflagrating spoon and heat it.

When the sulphur catches fire, introduce the deflagrating spoon into one of the jars of oxygen and allow the sulphur to burn. Keep the mouth of the jar closed. When the sulphur stops burning, remove the deflagrating spoon, allow the jar to cool and add 2 ml of water to the jar and shake. Test the solution with blue and red litmus papers and show that blue litmus paper turns red while red litmus paper remains unaffected.

Repeat the experiment with red phosphorus and show that the solution formed in this case is also acidic

Use the above observations to elicit from the pupils that

i. sulphur and phosphorus form oxides when they burn in oxygen,

ii. these oxides give acidic solutions.

Hold the magnesium ribbon with tongs, ignite it and introduce the burning ribbon into a jar of oxygen.

When magnesium stops burning, allow the jar to cool and add 2 ml of water to the jar and shake. Test the solution or suspension with blue and red litmus papers as before and show that blue litmus paper remains unchanged while red litmus paper turns blue.

Cut a small piece (pea size) of sodium, (handle sodium with forceps) and take it in a deflagrating spoon. Heat it and allow it to catch fire. Introduce the burning sodium into a jar of oxygen

When sodium stops burning, remove the deflagrating spoon and heat it again to burn off any unburnt sodium. (This check up is necessary before the spoon is dipped into water as sodium reacts vigorously with water). If no burning is noticed add 10 ml of water to the jar and dip the deflagrating spoon in it to make the residue react with water. Test the resulting solution with blue and red litmus papers as before and show that the solution in this case is basic.

Using the above observations, elecit from the pupils that

i metals like magnesium and sodium combine with oxygen forming oxides,

ii. these oxides react with water forming bases.

8.3 Properties of acids and bases

The observations made in experiment 29 may be made use of for establishing the properties of acids and bases. You may recall another important property of these substances, i.e., the electrical conductivity of their aqueous solutions (unit 4).

8.4 Neutralization

The pupils are now familiar with the characteristic properties of acids and bases Let them study in experiment 31 what happens to their properties when they are mixed with each other.

Experiment 31 Investigating the reaction between acid and alkali solutions

Time required : 40 minutes

Postlab discussion

Impress upon the pupils that when an alkali is added to an acid the acidic character of acid is lost and when an acid is added to alkali, the alkaline character of the alkali is lost. When the two, acid and alkali are mixed in equivalent proportion a neutral substance is formed. This process is called neutralization.

Questions and answers

Part I

1. What happens to the acid character of a solution when alkali is gradually added?
   The acidic character slowly decreases.
2. What happens to the alkaline character of a solution when acid is gradually added?
   The alkaline character gradually decreases.

Part II

What are the substances formed when hydrochloric acid reacts with sodium hydroxide? Write the equation for the reaction.

Sodium chloride and water are formed.

$HCl + NaOH \longrightarrow NaCl + H_2O$

You may then cite some of the neutralization processes we come across in daily life.

8.5 Salts

The pupils have now learnt that a salt is obtained when an acid and a base react. However, this is not the only method by which a salt is produced. To make the pupils realise that salt can be made by other methods, let them perform experiment 32, which describes the preparation of copper sulphate by the action of dilute sulphuric acid on copper oxide.

Experiment 32 Preparation of a salt by the action of an acid on the oxide of a metal

Time required : 40 minutes

It is desirable to go round the class when the pupils start evaporating the solution. They may experience some difficulty in judging the crystallisation point.

Question and answer

Write the equation for the reaction between cupric oxide and sulphuric acid.

$CuO + H_2SO_4 \longrightarrow CuSO_4 + H_2O$

Experiment 33 describes the preparation of ferrous sulphate by the action of dilute sulphuric acid on iron.

Experiment 33 Preparation of a salt by the action of an acid on a metal

Time required : 40 minutes

Questions and answers

1. Write the equation for the reaction between iron and sulphuric acid.
   $Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2\uparrow$
2. What type of reaction is it?
   It is a displacement reaction in which hydrogen is displaced from sulphuric acid by the metal, iron.

BACKGROUND INFORMATION

In the experimental approach to acids and bases, these substances have been defined in terms of their important characteristics. An important question "why do all acids show similar properties"? still remains unanswered. Evidently there is some factor common to all of them.

Aqueous solutions of acids produce hydrogen gas when metals like iron, magnesium and zinc react with them, indicating thereby that acids contain hydrogen. Since their solutions conduct electricity, it is evident that aqueous solutions of acids contain ions. Hydrogen present in the acids may, therefore, be in the form of ions in their aqueous solutions. It is these hydrogen ions (H^+) and not molecular hydrogen that account for the acidic properties. This is clear from the following. Dry HCl gas does not turn blue litmus paper red. But it does turn blue litmus paper red when dissolved in water. In water, HCl ionises into H^+ and Cl^- ions; the hydrogen ions, H^+, are responsible for acidic properties They do not exist free but are hydrated, forming hydronium ions, H_3O^+.

Ionisation of acids in aqueous solution may be represented as

$$HCl(g) + H_2O(l) \rightleftharpoons H_3O^+ + Cl^-(aq)$$

$$HNO_3(l) + H_2O(l) \rightleftharpoons H_3O^+ + NO_3^-(aq)$$

$$H_2SO_4(l) + H_2O(l) \rightleftharpoons H_3O^+ + HSO^{4-}(aq)$$

$$H_2CO_3(l) + H_2O(l) \rightleftharpoons H_3O^+ + HCO_3^-(aq)$$

These ideas were first put forward by Arrhenius. According to him, an acid is a substance that gives hydrogen ions when dissolved in water. As per this definition, an acid need not necessarily be a molecule. Anions like HSO_4^- and HCO_3^- which give hydrogen ions, H^+, may also be called acids.

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+ + SO^{4--}(aq)$$

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+ + SO_4^{--}(aq)$$

All bases have common properties. There must also be a factor common to all of them. Like acids, aqueous solutions of bases conduct electricity. This suggests that they form ions when dissolved in water. A base reacts with an acid forming a salt and water. In this reaction, the identifying properties (except electrical conductance) of both acids and bases are lost.

Since hydrogen ion, H^+, accounts for the properties of acids, the above reaction suggests that the base must be producing an ion which removes the hydrogen ion, H^+. In aqueous solutions, the base sodium hydroxide, NaOH, contains sodium ions, Na^+ and hydroxyl ions, OH^-.

Ionisation of bases in aqueous solutions may be represented as

$$NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$$

$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$Ca(OH)_2(aq) \rightleftharpoons Ca^{++}(aq) + 2OH^-(aq)$$

The OH^- ion from the base reacts with the H_3O^+ ion from the acid to form water

$$H_3O^+ + OH^-(aq) \rightleftharpoons 2H_2O$$

It may thus be postulated that the ion common to all bases is the hydroxyl ion, OH^-

A base is, therefore, a substance which forms hydroxyl ions, OH^-, in aqueous solutions.

Explanation of the properties of acids and bases in terms of Arrhenius concept

a. Electrical conductance of acids and bases is due to the existence of ions (anions and cations) in their aqueous solutions.

b Neutralization of an acid by a base and vice-versa may be explained in terms of the following ionic reactions·

$$HCl(g) + H_2O(l) \rightleftharpoons H_3O^+ + Cl^-(aq)$$

$$NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$$

$$H_3O^+ + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + 2H_2O$$

The net reaction involved in an acid-base neutralization may, therefore, be

$$H_3O^+ + OH^-(aq) \rightleftharpoons 2H_2O$$

c. Formation of hydrogen gas when acids react with certain metals may be explained as follows .

$$Zn(s) \longrightarrow Zn^{++}(aq) + 2e$$

$$2\,HCl + 2H_2O \longrightarrow 2\,H_3O^{+} + 2\,Cl^{-}\,(aq)$$

$$2\,H_3O^{+} + 2e \longrightarrow H_2\uparrow + 2H_2O$$

d. Decomposition of carbonates upon reaction with acids, liberating carbon dioxide gas may be explained as shown below :

$$Na_2CO_3(aq) \longrightarrow 2\,Na^{+}\,(aq) + CO_3^{--}(aq)$$

$$2\,HCl\,(g) + 2H_2O\,(l) \longrightarrow 2H_3O^{+} + 2\,Cl^{-}\,(aq)$$

$$CO_3^{-}\,(aq)^{-} + H_2O\,(l) \longrightarrow HCO_3^{-}\,(aq) + OH^{-}\,(aq)$$

$$H_2CO_3^{-}\,(aq) \longrightarrow CO_2\uparrow(aq) + OH^{-}\,(aq)$$

$$2\,H_3O^{+} + 2\,OH^{-}\,(aq) \longrightarrow 4\,H_2O$$

Strength of an acid or a base

An acid produces hydrogen ions in aqueous solutions. $\overset{+}{H}$ ions are responsible for acidic character. The strength of the acid therefore, depends upon the number of hydrogen ions produced in the solution. This in turn depends upon the degree of ionisation of the acid. The larger the degree of ionisation, the stronger is the acid. Degree of ionisation of HCl is quite high and that of acetic acid is very low. Hydrochloric acid is strong and acetic acid weak.

A base produces hydroxyl ions in aqueous solution. The degree of ionisation of a base determines its strength. The larger the degree of ionisation, the stronger is the base. Sodium hydroxide, NaOH, and potassium hydroxide, KOH, are strong bases since they have a high degree of ionisation. Calcium hydroxide, $Ca(OH)_2$, and ammonium hydroxide, NH_4OH, are regarded as weak bases since their degree of ionisation is low.

Quiz

1. Hydrogen chloride is a gaseous compound of hydrogen and chlorine. It does not conduct electricity when it is liquefied or when it is a gas or when it is dissolved in benzene, but it does conduct electricity when it is dissolved in water. Hydrogen chloride is more soluble in water than in benzene. Which of the following best fits these facts ?
   - i. Gaseous hydrogen chloride consists of widely separated ions
   - ii. Liquid hydrogen chloride consists of ions packed closely together
   - iii. Gaseous hydrogen chloride consists of widely separated molecules
   - iv. A solution of hydrogen chloride in toluene consists of widely separated hydrogen and chloride ions.
2. Given two white powders, one a basic oxide and the other a carbonate, describe how by applying the same test to each of the two substances, you could determine the nature of each.
3. Name the compound which is
   - a. often used to neutralize excess acid in the stomach
   - b. a salt used as baking powder
   - c. used in curing acidity of the soil
   - d. a gas that reacts with ammonia to form white fumes.
4. Write down equations for the following reactions
   - a. Hydrochloric acid + Sodium hydroxide
   - b Sodium hydroxide + Nitric acid
   - c Sodium hydroxide + Sulphuric acid
   - d. Lime water + Nitric acid.
5. Tick off the correct answer in the following.
   - a Magnesium oxide reacts with water forming a solution which
     - i. turns blue litmus paper red
     - ii. turns red litmus paper blue
     - iii. does not affect either.
   - b. Hydrochloric acid reacts with magnesium liberating
     - i. oxygen
     - ii. nitrogen
     - iii. hydrogen
     - iv. carbon dioxide.
   - c. An oxide of an element on dissolving in water forms an alkaline solution. Which of the following elements could it be ?
     - i. carbon
     - ii. sulphur
     - iii. sodium
     - iv. phosphorus.

CHEMICALS AND EQUIPMENT FOR THE LABORATORY EXPERIMENTS

The following lists serve as a guide for the purchase of chemicals and equipment required for one academic year. Chemicals are consumed and stocks are to be replenished. Equipment consists of items which can be reused, such as glass apparatus and hardware. It may not be possible to supply some items of the equipment individually to each pupil. In such cases, only a few may be stocked and used by the pupils by turns.

Chemicals

The following list indicates the amounts required for a class of 40 pupils, each pupil working individually as far as possible.

Substance	Quantity	Used in experiment
Aluminium sheets (2 cm. sq)	40 pieces	27
Aluminium wire (16 gauge)	200 g	27
Alcohol (for dissolving phenolphthalein)	100 ml	29
Ammonium chloride	100 g	18, 21
Barium chloride	25 g	18
Blue litmus paper	50 books	24, 29, 30
Calcium oxide	300 g	17, 18, 22, 25
Candle wax	100 g	27
Charcoal blocks with cavity	40	28
Charcoal (lumps)	80	25
Charcoal powder	50 g	25, 28
Copper carbonate	50 g	17
Copper foil (2 cm. sq)	40 pieces	27
Copper sulphate	50 g	17, 18
Copper wire (16 gauge)	200 g	27
Cupric oxide	250 g	25, 32
Hydrochloric acid (conc.)	2 litres	19, 22, 24, 29, 31
Iron nails (5 cm)	40	17
Iron powder	250 g	33
Iron strips (2 cm. sq)	40 pieces	27
Iron wire (16 gauge)	200 g	27
Lead oxide	25 g	28
Lead shot	50	27
Lead wire (16 gauge)	200 g	27
Litmus powder	2 g	31

Substance	Quantity	Used in experiment
Magnesium ribbon	2 rolls	17, 27, 29
Manganese sulphate	25 g	19
Marble chips	40 g	22
Mercuric oxide	25 g	17
Methyl orange	2 g	29
Oxalic acid	5 g	19
Phenolphthalein	2 g	29
Potassium permanganate	125 g	19, 20, 24
Red litmus paper	50 books	29, 30
Sodium carbonate	125 g	18, 29
Sodium hydroxide	25 g	31
Sodium nitrite	50 g	21
Sodium sulphate	20 g	18
Splinters	320	Frequent use
Spirit	10 litres	Frequent use
Sulphur	100 g	26
Sulphuric acid (conc.)	1 litre	23, 29, 32, 33
Tin granules	50 g	27
Turmeric paper	50 books	29
Zinc dust	50 g	19
Zinc granules	300 g	19, 23, 27

Equipment	Quantity	Used in experiment
Battery (2 volts)	40 sets	27
Beakers (100 ml)	40	25, 31
Beakers (250 ml)	40	32, 33
Blocks of iron (5 cm. cube)	40	27
Box of matches	40	Frequent use
Blow pipes	40	28
Connecting wire (20 gauge)	1 Kg	27
Corks (assorted)	1 gross	20. 21, 22, 23
Droppers	80	29, 31
Evaporating dishes	40	31, 32, 33
Filter paper circles	50	33
Funnels	40	32, 33
Funnel stands	40	32, 33
Glass rods	40	36, 31, 32, 33
Graduated test tubes	40	18, 19, 21, 23, 29

Equipment	**Quantity**	**Used in experiment**
Iron hammers	40	27
Iron stands (with adjustable clamps)	40	20, 21, 22, 23, 24
Kerosene lamps	40	28
Metal discs (10 cm. diameter, carrying wires of Al, Cu, Fe and Pb, 16 gauge and 10 cm. long each)	40	27
Pair of tongs	40	17, 28, 32, 33
Pen knives	40	27
Rubber stoppers (carrying delivery tubes)	40	17, 20, 21, 22, 23, 24, 25, 29
Sand bath	40	31, 32
Sand paper (5 cm. sq)	40	27
Spirit lamps	40	Frequent use
Tap keys	40	37
Test tubes (hard glass)	50	18, 20, 21, 22, 23, 24, 25
Test tubes (ordinary)	200	Frequent use
Test tube brushes	40	Frequent use
Test tube holders	40	17, 18, 19, 25
Test tube stands	40	Frequent use
Thread (cotton)	1 reel	17
Torch bulbs (1.5 or 2.0 volts)	40	27
Tripods	40	32, 33
Troughs	40	20, 21, 23
Watch glasses	160	17, 18, 30
Water bath	40	33
Wire gauze	40	25, 32, 33

The following items may be brought by the puplis on the day of the experiment, or may be supplied by the teacher.

Copper vessels
Curd
Lemon
Milk
Soil samples from different places
Tamarind
Toothpaste
Washing soap

EQUIPMENT REQUIRED FOR DEMONSTRATION EXPERIMENTS

The following is the list of apparatus and chemicals needed for showing demonstration experiments of units 5-8, *for one section only*. The chemicals which are consumables have to be replenished if the experiments are to be demonstrated to another section. The apparatus, however, could be used any number of times.

Apparatus	**Quantity**
Aluminium sheet 1sq. cm	1
Asbestos square (15 cm × 15 cm)	1
Beakers (150 ml)	3
Beaker (250 ml)	1
Beehive shelf	1
Blow pipe	1
Bulb 2.5 V	1
Bulb holder	1
Clay pipe triangles	2
Copper strip (12 cm × 12 cm × 0.2 cm)	1
Conical flasks (150 ml)	4
Corks(assorted)	6
Cork borer set	1
Connecting wire	1 meter
Crocodile clips	2
Crucibles	2
Cycle pump	1
Deflagrating spoon	1
Delivery tubes	2
Droppers	2
Evaporating dish	1
Flask R. B.(250 ml)	1
Funnels	2
Funnel stand	1
Gas jars with covers	6
Glass plate (10 cm × 10 cm)	1
Glass rods	2
Hammer	1
Iron stand with clamps	1
Iron strips (5 cm × 1 cm)	2
Kerosene lamp	1
Measuring cylinder (25 ml)	1
Mortar and pestle	1
Pen knife	1
Reagent bottle (4 oz)	1
Rubber stoppers (assorted)	6
Spirit lamps	2
Test tubes	4
Test tube (hard glass)	1
Test tube holder	1
Thermometer	1

Apparatus	**Quantity**
Thistle funnel	1
Tongs	1
Tripods	2
Trough	1
Watch glass	1
Wire gauze	1

CHEMICALS

Aluminium powder	1 g
Ammonium chloride	10 g
Animal charcoal	1 g
Calcium oxide	10 g
Candle	1
Carbon disulphide	5 ml
Charcoal pieces	4
Charcoal powder	3 g
Copper powder	2 g
Copper sulphate	2 g
Copper turnings	4 g
Copper wire (20 gauge)	10 cm
Cupric oxide	2 g
Filter papers circles	10
Glycerine	5 ml
Hydrochloric acid (conc.)	100 ml
Iron filings	1 g
Iron wire (20 gauge)	5 cm
Lead metal	4 g
Lead carbonate	1 g
Lead nitrate	1 g
Litmus paper (Blue)	1 book
Litmus Paper (Red)	1 book
Litmus powder	5 g
Magnesium ribbon	50 cm
Manganese dioxide	8 g
Marble chips	25 g
Match boxes	6
Mercury	5 ml
Nitric acid (conc.)	15 ml
Ores of different metals	1 set
Oxalic acid	1 g
Phosphorous	1 g
Potassium bromide	1 g
Potassium chlorate	12 g
Potassium permanganate	3 g
Sand paper	1 sheet
Silver nitrate	1 g
Spirit	2 litres
Splinters	20
Sodium metal	3 g
Sodium nitrite	10 g
Sugar	2 g
Sulphur powder	100 g
Sulphuric acid (conc.)	50 ml
Tin metal	6 g
Zinc granules	40 g

International Atomic Weights

Element	Symbol	Atomic No.	Atomic Wt.	Element	Symbol	Atomic No.	Atomic Wt.
Actinium	Ac	89	227	Einstenium	Es	99	[254]
Aluminium	Al	13	26 98	Erbium	Er	68	167 27
Americium	Am	95	[243]*	Europium	Eu	63	152.0
Antimony	sb	51	121 76	Fermium	Fm	100	[253]
Argon	Ar	18	39 994	Fluorine	F	9	19.00
Arsenic	As	33	74·91	Francium	Fr	87	[223]
Astatine	At	85	[210]	Gadolinium	Gd	64	157 26
Barium	Ba	56	137 36	Gallium	Ga	31	69.72
Berkelium	Bk	97	[249]	Germanium	Ge	32	72 60
Beryllium	Be	4	9.013	Gold	Au	79	197 0
Bismuth	Bi	83	209 00	Hafnium	Hf	72	178.50
Boron	B	5	10.82	Helium	He	2	4 003
Bromine	Br	35	79.916	Holmium	Ho	67	164.94
Cadmium	Cd	48	112.41	Hydrogen	H	1	1·008
Calcium	Ca	20	40 08	Indium	In	49	114·82
Californium	Cf	98	[251]	Iodine	I	53	126·91
Carbon	C	6	12 011	Iridium	Ir	77	192 2
Cerium	Ce	58	140.13	Iron	Fe	26	55.85
Cesium	Cs	55	132·91	Krypton	Kr	36	83.80
Chlorine	Cl	17	35 457	Lanthanum	La	57	138 92
Chromium	Cr	24	52 01	Lead	Pb	82	207.21
Cobalt	Co	27	58.94	Lithium	Li	3	6 940
Copper	Cu	29	63·54	Lutetium	Lu	71	174 99
Curium	Cm	96	[247]	Magnesium	Mg	12	24 32
Dysprosium	Dy	66	162 51	Manganese	Mn	25	54 94

Element	Symbol	Atomic No	Atomic Wt.	Element	Symbol	Atomic No	Atomic Wt
Mendelevium	Md	101	[256]	Ruthenium	Ru	44	101.1
Mercury	Hg	80	200.61	Samarium	Sm	62	150.35
Molybdenum	Mo	42	95.95	Scandium	Sc	21	44.96
Neodymium	Nd	60	144.27	Selenium	Se	34	78.96
Neon	Ne	10	20.183	Silicon	Si	14	28.09
Neptunium	Np	93	[237]	Silver	Ag	47	107.880
Nickel	Ni	28	58.71	Sodium	Na	11	22.991
Niobium	Nb	41	92.91	Strontium	Sr	38	87.63
(Columbium)				Sulphur	S	16	32.066
Nitrogen	N	7	14.008	Tantalum	Ta	73	180.95
Osmium	Os	76	190.2	Technetium	Tc	43	[99]
Oxygen	O	8	16	Tellurium	Te	52	127.61
Palladium	Pd	46	106.4	Terbium	Tb	65	158.93
Phosphorus	P	15	30.975	Thallium	Tl	81	204.39
Platinum	Pt	78	195.09	Thorium	Th	90	232.05
Plutonium	Pu	94	[242]	Thulium	Tm	69	168.94
Polonium	Po	84	210	Tin	Sn	50	118.70
Potassium	K	19	39.100	Titanium	Ti	22	47.90
Praseodymium	Pr	59	140.92	Tungsten	W	74	183.86
Promethium	Pm	61	[145]	Uranium	U	92	238.07
Protactinium	Pa	91	231	Vanadium	V	23	50.95
Radium	Ra	88	226.05	Xenon	Xe	54	131.30
Radon	Rn	86	222	Ytterbium	Yb	70	173.04
Rhenium	Re	75	186.22	Yttrium	Y	39	88.92
Rhodium	Rh	45	102.91	Zinc	Zn	30	65.38
Rubidium	Rb	37	85.48	Zirconium	Zr	40	91.22

•Values in brackets are mass numbrs of longest-lived or best-known isotopes.

Dear Reader

The present experimental editions have been brought out as a result of the Education Commission's recommendation that chemistry should be taught as a separate discipline in our schools from class 6 onwards. This has become necessary in view of the great strides that chemistry has made in recent years and the consequent need for the pupils in our schools to be exposed to the modern ideas and trends in the subject This modest effort, it is hoped, will fulfil that long felt need. It is aimed at presenting the subject in a simple language so as to be comprehensible to pupils in the age group 10 plus to 12 plus

The emphasis is on the pupils' learning the basic concepts of chemistry through experimentation and observation Stress is laid on enabling the pupils to develop sufficient skill in simple laboratory techniques and help them appreciate the need for careful observation. These have been the guiding principles for the write up of the laboratory manual and the text. The teacher's guide, as its name implies, is expected to fill in any gaps in the academic make up of the teacher to the extent required in presenting the material to the pupils.

The authors are conscious of the shortcomings and limitations of the material prepared by them. They are aware also that any meaningful improvement in the style and the contents of the books will be possible only after a candid and critical comment on the material, by all those who are genuinely interested in the improvement of standard of chemistry in our schools. The members of the Study Groups have, therefore, great pleasure in sending you copies of the books and invite your comments and constructive suggestions, by way of answers to the questionnaire after a survey of the books

Any useful suggestions beyond the questionnaire are welcome.

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Convener,
Chemistry Study Groups.

QUESTIONNAIRE

On the school chemistry laboratory manual, text book and teacher's guide for classes 6 & 7

N. B. :—*In case the space provided against each question is insufficient, please attach additional papers. Specific and concrete suggestions far any modification, alteration and improvement of the material will be appreciated.*

1. Are the experiments in the laboratory manual simple enough to be easily done by the pupils ?

   If not, which experiments, do you feel are difficult ?

   How would you like them to be modified ?

2. Are the experiments suited to arouse the curiosity and interest of the pupils ?

   Do they lead to an appreciation of the basic principles of chemistry ?

   If not, what improvements do you suggest ?

3. In general, are the experiments in the laboratory manuals well planned ?

4. Do you think that the experiments mentioned in the manuals are adequate ?

   If not, what additional experiments would you like to be introduced and at what places ?

5. Do you think that the materials required for the experiments would be easily available ?

   If not, can you suggest alternatives, using more easily available materials, to bring out the same ideas ?

6. What, in your opinion, is the scope and extent of improvisation that can and may be made to replace the equipment mentioned ?

7. What, in the context of the present condition of our schools, is your opinion about the feasibitity of the idea of pupil experimentation ?

8. The material in the text has been prepared, to a large extent, to help the pupil learn chemistry by drawing conclusions from his own observations in the laboratory.

   Do you think that this objective has been broadly realised ?

9. Is the language on the whole simple enough for an easy understanding of the subject ?

   If not, point out where the language could be simplified.

10. Do you think that the material as developed in the text makes for continued and logical development of ideas and concepts ?

11. Are the contents of each unit within the comprehension of pupils of the age group of 10+ and 12+ ?

    If not, is it possible to express the same concept in a simpler way ?

12. Do you think that the sequence of the units is satisfactory ? If not, Please suggest how the sequence may be rearranged to enable a better development of the subject.

    Give reasons for your rearrangement

13 Are the diagrams and illustrations suggestive and self-explanatory ? State how they may be improved. Do you think that the present illustrations are adequate ?

    If not, which additional diagrams do you suggest ?

14. Do the demonstration experiments mentioned in the teacher's guide bring about the concepts ?

    If not, what alternative experiments would you suggest ?

15. The teacher's guide is intended to enable the teacher to teach the subject confidently and more effectively. It guides him to plan the lesson and provides background information to extend the sphere of his knowledge and answer questions which might be put by intelligent pupils. How far does this teacher's guide serve the purpose for which it is intended ?

16. Do you think that the teacher's guide contains sufficient information for the teacher to present the topic effectively ?

    If not, what additional information may be given and where ?

17. The aim of this effort is to introduce the pupil to modern concepts in chemistry and promote a sense of experimentation and inquiry in him. Do you think that this would be achieved by the text and laboratory work presented ?

    What alterations would you suggest in this regard ?

18. What other pertinent comments and suggestions would you like to make ?

Signature

Station : Name :

Date : Address :